1958

The comparison of inductive and deductive methods of teaching high school chemistry

O'Connell, Sister Ernestine Marie

Boston University

http://hdl.handle.net/2144/6240

Boston University
BOSTON UNIVERSITY
SCHOOL OF EDUCATION

Thesis

THE COMPARISON OF INDUCTIVE AND DEDUCTIVE METHODS OF TEACHING
HIGH SCHOOL CHEMISTRY

Submitted by

Sister Ernestine Marie O'Connell, S.C.H.
(A.B., Boston University, 1945)
(A.M., Boston University, 1946)

In Partial Fulfillment of Requirements for
the Degree of Doctor of Education

1958

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CHAPTER I
CHARACTERISTICS OF THE STUDY

1. Introduction

Purpose.—The comparison of inductive and deductive methods in teaching secondary school chemistry is being studied. Trials of a full year program and of a single unit have been made, and data have been correlated in both studies. Phase I of this paper deals with the treatment of data derived from a full year's experiment in three schools; Phase II treats of forty schools testing one unit on equation-balancing.

The data obtained from employing inductive and deductive methods will reveal whether there is any significant statistical difference in achievement caused by method variation. The difference may be judged tangibly on the full year course by the results on standardized chemistry tests, viz., the Anderson Chemistry Test and the Cooperative Chemistry Test. On the equation-balancing unit the measure will be the difference in performance on the symbolic unit test (Appendix F) administered before and after the unit is taught.

The writer hopes that the challenge of the inductive method may


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incite a greater interest in the true study of chemistry, diminish
the number of high school "memorizers", and bring forth a group of
vitally interested students who, having their scientific curiosity
thoroughly aroused, will go on with scientific pursuits. 1/

"Any good student is irked to death by the memorization
and 'plugging into formulas' and routineness with which science
is taught these days .... One way to get more science students
would be to set up courses in a way to encourage independent
investigations by students .... but too few teachers are them-
seves trained in investigative methods."

Scope.-- One girls' school in Dorchester, Massachusetts, averaging
thirty-five pupils per class, was the basic testing area. In this
school four classes were taught inductively and four classes were
taught deductively. On these classes complete data of capacity and
achievement were collected. In addition, two other schools of mixed
classes of 35-40 students each were used. These classes were provided
with inductive laboratory guide sheets and a course of study prepared
by the writer. After having completed the course presented in this
manner, these students were given the Anderson and Cooperative standard-
ized chemistry tests and the California Test of Mental Maturity. 2/
Records from these schools were also used to supply data on preceding
classes taught by the same teacher, but in the conventional manner.

1/E. J. Stennerson, "Editorial", Chemical and Engineering News,
(March 21, 1939), 35:1215.

2/California Test of Mental Maturity, advanced form, California Test
Bureau, 5915 Hollywood Boulevard, Los Angeles 28, California, 1946.
This was done to determine whether the figures obtained from the Dorchester test school were exaggerated because of the teacher, or whether the same material, in the hands of other trained teachers would give comparable results. The schools selected for these studies were ones in which the present chemistry teachers had been located for the four previous years, so that there would be a minimum of variation in the teaching procedures.

2. Tools and Procedures

Syllabus. -- The writer worked in her own school to develop and test these theories and techniques. For standard, deductive, macro work, the basic course of study of the Archdiocese of Boston was used, (Appendix A) along with the College Entrance Book Company Laboratory manual and workbook. 1/

For inductive teaching, a new course of study was constructed by the writer, and a set of laboratory guide sheets constructed to accompany it. (Appendix B and C) A letter was sent to the colleges and nursing schools most frequently accepting students from the school population under consideration. (Table 1 lists these schools.) In the letter the professors of freshman chemistry were asked to list the topics they considered of greatest importance at the secondary level.

Table 1. Colleges and Nursing Schools Polled in Syllabus Search

<table>
<thead>
<tr>
<th>School Number</th>
<th>School</th>
<th>Location</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>St. Anselm's College</td>
<td>Manchester, New Hampshire</td>
</tr>
<tr>
<td>2</td>
<td>Boston City Hospital</td>
<td>Boston, Massachusetts</td>
</tr>
<tr>
<td>3</td>
<td>Boston College</td>
<td>Chestnut Hill, Massachusetts</td>
</tr>
<tr>
<td>4</td>
<td>Boston University</td>
<td>Boston, Massachusetts</td>
</tr>
<tr>
<td>5</td>
<td>Emmanuel College</td>
<td>Boston, Massachusetts</td>
</tr>
<tr>
<td>6</td>
<td>Fairfield University</td>
<td>Fairfield, Connecticut</td>
</tr>
<tr>
<td>7</td>
<td>Holy Cross College</td>
<td>Worcester, Massachusetts</td>
</tr>
<tr>
<td>8</td>
<td>State Teachers College</td>
<td>Framingham, Massachusetts</td>
</tr>
<tr>
<td>9</td>
<td>Mass. General Hospital</td>
<td>Boston, Massachusetts</td>
</tr>
<tr>
<td>10</td>
<td>Northeastern University</td>
<td>Boston, Massachusetts</td>
</tr>
<tr>
<td>11</td>
<td>Regis College</td>
<td>Weston, Massachusetts</td>
</tr>
<tr>
<td>12</td>
<td>St. Elizabeth's Hospital</td>
<td>Brighton, Massachusetts</td>
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<td>13</td>
<td>St. Joseph's College</td>
<td>Hartford, Connecticut</td>
</tr>
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<td>14</td>
<td>Simmons College</td>
<td>Boston, Massachusetts</td>
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<td>15</td>
<td>Stonehill College</td>
<td>Northeastern, Massachusetts</td>
</tr>
<tr>
<td>16</td>
<td>Tufts College</td>
<td>Medford, Massachusetts</td>
</tr>
</tbody>
</table>

The topics considered of greatest importance by these professors who responded to the poll are listed in Table 2. Once this material was collected, it was studied in conjunction with the major courses of study available, including the Philadelphia chemistry syllabus of 1950; the New York State Regents' requirements; the College Entrance Board suggested syllabus for 1940; the New England Association of
Chemistry Teachers' syllabus for 1940; the Catholic University of America's syllabus for affiliated schools, 1943 revision; and the American Chemical Society's suggestions for 1943.

Table 2. Key Topics Reported in Curriculum Survey

<table>
<thead>
<tr>
<th>Number</th>
<th>Topic</th>
<th>Schools Reporting</th>
<th>Total Schools</th>
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<tbody>
<tr>
<td>1</td>
<td>Valence</td>
<td>4,5,7,9,11,13,14,16</td>
<td>9</td>
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<tr>
<td>2</td>
<td>Simple equations and laws</td>
<td>4,5,7,9,10,11,13,14,16</td>
<td>9</td>
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<tr>
<td>3</td>
<td>Problems involving weights</td>
<td>1,2,4,5,9,11,12,14,16</td>
<td>9</td>
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<td>4</td>
<td>Atomic theory and periodic classification</td>
<td>1,3,4,5,7,10,11,14</td>
<td>8</td>
</tr>
<tr>
<td>5</td>
<td>Acids, bases, and salts</td>
<td>2,4,5,9,11,13,16</td>
<td>7</td>
</tr>
<tr>
<td>6</td>
<td>Ionization</td>
<td>2,4,5,9,11,13,16</td>
<td>7</td>
</tr>
<tr>
<td>7</td>
<td>Solutions and suspensions</td>
<td>2,4,5,9,12,14</td>
<td>6</td>
</tr>
<tr>
<td>8</td>
<td>Oxidation-reduction</td>
<td>5,11,13,14,15,16</td>
<td>6</td>
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<tr>
<td>9</td>
<td>Gas laws and kinetic theory</td>
<td>1,3,4,6,7,15</td>
<td>6</td>
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<td>10</td>
<td>Descriptive material limited to popular elements</td>
<td>2,9,11,12,14,16</td>
<td>6</td>
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<td>11</td>
<td>Coverage of any good textbook</td>
<td>3,6,7</td>
<td>3</td>
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<td>12</td>
<td>Non-mathematical radioactivity</td>
<td>16</td>
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</table>

The final course to be used consisted of nine topics:

1. Atomic structure and nomenclature (Topics 1 and 2, chart 2)
2. Periodic classification (Topic 4)
3. Equation balancing and laws (Topic 2)
4. Stoichiometric problems (Topic 3)
5. Solutions and suspensions (Topic 7)
6. Ionization—acids, bases and salts (Topics 5 and 6)

* School numbers refer to Table 1.
7. Gas Laws
8. Oxidation-reduction
9. Radioactivity and nucleons

Information obtained. — For each student in the test classes the following data were listed:

1. Intelligence test scores (California, advanced)
2. Anderson Chemistry Test scores (administered in April)
3. Cooperative Chemistry Test scores (administered in May)

Definition of terms. — The designations "deductive" and "inductive" may here need some clarification. Deduction is used broadly in contrast to induction as meaning a reasoning from the general to the particular, from the implicit to the explicit, as contrasted with reasoning from particular facts to general truths, or from a part to a whole.

According to the New Century Dictionary, 1/

"The deductive method is the method of reasoning by which, from assumed or established general principles, concrete applications or conclusions are deduced."

Induction, on the other hand, is an act of reasoning from a part to a whole, from particulars to generals, or from the individual to the universal. 2/

Aristotle considered induction a form of reasoning, which, when perfect, consisted in a complete enumeration of all the particulars contained in the inferred generalization.

Bacon introduced the inductive method which consists in the

inferring that what has been observed or established in respect to a
part, individual or species, may, on the ground of analogy, be affirmed
or received of the whole to which it belongs. Such inference ascends
from parts to wholes and forms, from the general analogy of nature,
or the special presumptions of the case, conclusions which have a
greater or less degree of force and which may be strengthened or
weakened by subsequent experience or experiment, but which, in the long
run, will rectify themselves by reason of repeated operations. 1

In Webster 2 we find induction applied to mathematics:

"The process of demonstration in which the general
validity of a law is inferred from its observed validity
in particular cases by proving that if the law holds in
a certain case it must hold in the next case, and so in
the next, and so on endlessly."

The writer has in mind an inductive and a deductive method of
laboratory procedure which are in direct contrast one with the other.

By deductive chemistry is meant that traditional high school
course in which theory and descriptive chemistry are taught in the
classroom and in which a manual of directions is followed by the
student in performing laboratory experiments for one or two hours a
week. In this way he sees for himself the physical phenomena dis-
cussed in the class.

By inductive chemistry, on the other hand, is meant that type
of course in which theory and theory only is pronounced in the class-
room. However, the course does not violate the New England Association

1 Roger Bacon, Opus Majus, Edited by L.H. Bridges, Oxford University
Press, 1897.
2 Merriam Webster, op. cit.
of Chemistry Teachers' Syllabus which is generally accepted as adequate in the New England area. (Appendix E) In the laboratory the student discovers for himself the physical and chemical properties of elements and compounds, having no formal set of directions, but only those he has assembled or has discovered by perusing chemical literature.

3. Design of the Experiment

Experimental group. — The girls' high school in Dorchester, Massachusetts, previously mentioned, was used for the general testing. The students were from an urban area of moderate to low income brackets. The college preparatory section was somewhat screened so that the chemistry classes had the better half of the junior year students.

Four consecutive classes in the school were taught chemistry by the traditional deductive method and complete records were kept of their capacity and achievement.

The next four classes received teaching on chemical principles and theory only and were left to themselves to devise experiments suitable for the discovery of the physical and chemical properties of assigned elements and compounds. This group is the one herein designated as the inductive group. For laboratory work these inductive classes were given a General Outline of Elements and Compounds. This was a sort of guide to keep the students within bounds in experimentation, as it was found that many high school students lacked that discipline of mind necessary to choose only pertinent experiments.
If a student finished his assigned work early, he was allowed to test and experiment as he pleased, but first he was required to do the laboratory work necessary to give him positive information concerning the facts in the basic guide sheet outline. This outline follows:

**GENERAL OUTLINE FOR ELEMENTS AND COMPOUNDS**

A. Occurrence and methods of extraction

B. Preparation
   1. Laboratory (description and equations)
   2. Commercial (description and equations)

C. Properties
   1. Physical
      a. Density
      b. Color
      c. Odor
      d. Taste
      e. Solubility
      f. Change of state
      g. Special (allotropic forms, isotopes, etc.)
   2. Chemical
      a. Stability
      b. Reaction to burning
      c. Reaction with elements (description and equations)
      d. Reactions with compounds (description and equations)
      e. Special (reducing or oxidizing agent; acid or base)
      f. Test for the element or compound (equations)

D. Structure
   1. Atomic diagram
   2. Molecular diagram
   3. Ionic diagram

E. Uses

F. Equations not covered in the preceding sections

G. Diagram of the apparatus used for the laboratory preparation.
Outside schools.--Two other schools taught using the inductive syllabus prepared by the writer. The data they obtained from a year's trial of the inductive method were relayed to the writer. School B, located in the West Indies, had a population from the moderate to upper income groups, all white, mixed male and female students. The other school, C, was in an urban center in Ohio and drew boys and girls from the middle bracket population. In each of these schools the chemistry course was taught in the eleventh grade. Records were sent of the preceding class taught in the traditional deductive manner by the same teacher in the previous year. These data were used as norms from which to judge the inductive class results. All students involved were from the college preparatory classes.

In Phase I there were twelve classes involved in all, each consisting of approximately thirty-five students, making a total of four hundred and twenty students, half in the control groups (deductive) and the other half in the experimental groups (inductive). Phase II treats with a larger number of classes and students on a single unit of the course.

In inductive laboratory learning the student picks up information as he progresses through his experiments. In studying the electrochemical series, for instance, he will discover, among other things, that small quantities of reagents can be handled more easily than large ones. By trial and error he has learned the value of clean glassware; he has gained facility in handling stock and reagent bottles; he has found that concentrated and dilute acids do not yield the same products in reactions with metals, and that concentrated acids are not
better and faster reactors, as he might have thought they were. His puzzles he can not answer all at once. The one mentioned here, for instance, will not be solved by him at the time he is testing the reactivity of metals, yet, by meeting the problem he has developed a mental set ready and receptive to later sections on ionization. In using nitric acid he has found it does not behave like other acids on metals such as lead or zinc. His burning splint test failed to reveal the identity of the gaseous product. He could not write all the equations for these reactions when he first met them, but he learned to refer to his text to find out the expected products and then to find tests for them. He developed a sense of mastery, and a respect for and a familiarity with his textbook. Further illustrations of the inductive laboratory learning that is to be expected may be found in Appendix F.
CHAPTER II
SURVEY OF RELATED LITERATURE

Despite the fact that, "...the very act of performing one experiment and permitting a class to conclude from it is unscientific practice," the battle between individual laboratory work and the lecture demonstration has waged with great intensity during the last two decades; and although the evident superiority of individual experimentation has long been recognized, the pressure of time, cost, and enlarged student enrollment in secondary schools has initiated trends that appear to jeopardize the place of the laboratory in the physical sciences, and especially in chemistry. Hodges, in an article published in the 1956 Journal of Chemical Education has remarked:

"The United States educational system is dedicated to the proposition that every child should have a broad educational experience in so far as he is able. To provide this type of education, and to be sure that our industrial economy is supplied with an adequate number of engineers and chemists, poses a problem with primarily rests with the school boards, principals, and teachers.

These days-- if they ever really existed-- when the top ten percent of the students were in the chemistry and physics classes are over. The American public school system is committed to the ideal of teaching all youth who enroll in them. As they attempt to educate all youth, all sorts of people are found in these classes-- dullards, young geniuses, future housewives, and future research scientists... To fail with this last group will result in an ever increasing shortage of personnel who make contributions that affect our way of life."


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She also deplores the practice of curriculum dilution which is gradually removing the subjects that the students call "hard" and is replacing them with new courses "designed to meet the needs of youth" because they will then be happy and have no need of study. A further malady, she points out, is the growing trend to pass all students in high school classes, regardless of the calibre of their work. 

"The paradox of modern science education seems to be that in order to give physics and chemistry to the gifted, all students must be admitted to the classes. This has troubled school officials for many years. They realize that this type of student is left pretty much to shift for himself while the teacher spends most of his time with the mediocre students so that he can pass them with a fairly clear conscience... The secondary school program should be so organized that the educational opportunities offered should match, utilize and challenge their abilities."

In 1900 only about ten percent of high school age youth were in high school; fifty years later close to ninety percent were in attendance. That ten percent of former days was a fairly select group coming from homes with good cultural and financial advantages; whereas today's ninety percent is a diversified group from all types of backgrounds and socio-economic conditions. Today's school is expected to take in minds of all calibres and do what it can with them. True, chemistry in many places is reserved for the college preparatory students or for the classical groups, where some selection may have been done. But, by and large, the students registering for junior and senior chemistry in

1/ Edith Hodges, op. cit., p. 406

our large urban high schools are not all of high or even average IQ.

Schlesinger, in 1935, defined the purpose of laboratory work as "training in observation, thought and action." The laboratory, he declared, was to be used by individual students to clarify and illustrate principles already discussed in the classroom; to give the students a feeling of the reality of science by actual contact with the phenomena the books describe; and to impress scientific facts on the mind of the student while he acquired manual dexterity and insight into basic scientific methods.

However, although discussions of inductive and deductive methods as such do not appear in the chemical educational literature of Schlesinger's time, his own paper indicates his awareness of the problem.

"The common demand that the laboratory manual must follow the order of presentation of the textbook and must contain many questions intended to review and clarify topics discussed in lecture or recitation indicates that the chief purpose of laboratory exercises is to support classroom work by illustration and example rather than to attain an independent objective."

It is to be noted that Schlesinger has used the terminology "the common demand" in the above quotation. Further study of his paper reveals that he is not wholly in sympathy with the "common ideas of his times.

"Any scheme of education is seriously deficient if it does not include planned training in the act of translating observation and thought into well considered action. Yet no part of


2/ Ibid., p. 54.

3/ Ibid., p. 57.
the duty of education has been more neglected.

In the synthesis of observation, reflection and action into a pattern of behavior, the laboratory work of the physical sciences is one of the most powerful tools. Moreover, properly designed laboratory exercises frequently demand of the student that he check his conclusions by further experimentation, not planned or laid out for him, but devised by himself. Next year another procedure will be tried; the students will be allowed to devise apparatus for carrying out such suggested reaction and will learn from experience what are the criteria of simple and convenient procedures.

The important point is that the student be allowed, in fact required, to put his suggestions into practice if it is at all feasible; and there is no reason why he should not try out a reaction that could not possibly yield the desired result, provided he can thereby discover for himself the reason for his failure."

An further, Schlesinger remarks that although the illustration of principles and the clarification of difficulties by direct contact with phenomena cannot be neglected, the main purpose must be achieved by selecting as the most significant part of laboratory work, exercises, the results of which the students cannot readily predict. These exercises, Schlesinger feels, should demand, as far as possible, some simple solution by experimental means. Success should be marked in terms of accuracy of observation and should be measured by increasing ability to use knowledge in drawing conclusions, and greater independence in planning tests to check the conclusions drawn. 1/

"To achieve this ideal in large classes without overstepping either the school's financial resources or the students' mental abilities, is a difficult task. It will require time, thought, imagination, and creative skill."

In 1944, Thelen of Chicago 2/ demonstrated experimentally that,

1/H. L. Schlesinger, op. cit., p. 75
by modification of existing procedures of instruction, a wider range of abilities in the area of scientific thinking could be learned without the sacrifice of subject matter content.

P.D. Curtis of Michigan states: 1/

"The two methods (inductive and deductive) are intrinsically complementary.... When the deductive method is supreme, as it is in the typical high school course, the assignments of materials to be studied in the textbook or other sources, precede the laboratory work.

Thus, with the inductive method, the laboratory work presents a problem or challenge which the deductive method rarely provides. If a pupil does not know an answer to a question, his curiosity is aroused; he is stimulated to try to find the correct answer."

As a further danger inherent in the deductive method, Curtis 2/ points out the following commonly experienced evils. When the pupil knows in advance what results to expect, he engages in one or more of the following practices:

1. He reports he has observed what he thinks he should have observed.
2. He changes data to make the answers 'come out right'.
3. He writes up an experiment without ever doing it.

The fundamental difference between the inductive and the deductive methods does not lie in the presence or absence of directions, but in the way in which the directions function. With the deductive method, specified procedures lead the pupil to results he expects; with the inductive, on the other hand, corresponding directions lead him to discover facts or principles previously unknown to him. 3/

2/Ibid., p. 223.
3/Ibid., p. 224.
It seems unquestionable that the inductive method provides better opportunities for teaching the elements of the scientific method. The inductive method makes the laboratory part of the course a fascinating adventure. The teacher who gives it a thorough trial will never discard it. Instead, he will make it a major supplement in all his science teaching.  

These ideas are corroborated by Himan 1 in his study of laboratory manuals in 1957. He enumerated eight leading evils of the workbook type of chemistry laboratory teaching:

1. Students are not required to write enough.
2. Attention is directed to the observations the author wants made.
3. Experiments are a hodgepodge of blanks, directions, and pictures.
4. Most directions implicitly give the answers to questions.
5. Multiple choice spoils critical thinking.
6. Ingenuity is never required.
7. The mathematical treatment of quantitative experiments in the same cookbook treatment as the qualitative.
8. No sketches of equipment are required.

Bunting says, 2 To develop the ability to think critically is one of the objectives of education. Science, by the very nature of its material and methodology would seem to offer fruitfulness toward the realization of this objective. 2 And adding to this, Elbert Weaver's statement that, "The ideal chemistry course would be an experience of learning by doing" 3 we have the underlying principles of the inductive method. True, the laboratory must teach safety, coordination of thought and action, training in observation of cause and effect; discernment of significant from irrelevant data; accurate reporting and manipulative

1/Fred Himan, Jr., "Laboratory Manuals and Workbooks in High School Science," Science Teacher (1943) 20:1116.
skill. Does the average high school chemistry teacher accomplish this? Otto says: 1/

"Any teacher will need the fundamental courses in general, qualitative, quantitative, organic, physical, colloidal, and instrumental chemistry in undergraduate work. He needs basic physics and enough mathematics to handle the above.... Sound training in knowing how to use technical literature.... Ability to prepare and write a thesis.... A good research attitude.... Chemistry is a science and teaching is an art."

Hall has noted: 2/

"Few secondary school teachers work from a self-prepared outline; they learn one published text.... Teaching as now generally practiced has lost its professional aspect of the past. Teachers are technicians--surface workers-- instead of true educators steeped in the knowledge that it pervaded by a background of philosophical understanding."

He went on to sympathize with the teacher who nowadays is swamped with ill-prepared students, credit seekers, and a whole generation of children who have never had the thrill of real accomplishment in the activities of intellectual discipline. He further argued that most school courses are compromises designed to reach the average student--still too hard for the poorer ones and too simple for those above average IQ--and hence are responsible for the tragic loss of the gifted. To remedy the situation he suggested the following changes in the curriculum: 2/

"Reduce detailed study of the metric system, gas laws, and density to such ideas as are directly related to chemical phenomena.... Leave out considerable material (the detailed study of the various members of the periodic table and related


3/ibid., p. 276.
compounds of the common elements). More time would be spent in the study of the Table and its use as a chemical tool. Traditional high school texts give great space to historical aspects and the high school students are not historically minded. Therefore, abandon the historical aspects of the study. Also eliminate consumer and industrial chemical items as major aspects. Increase coverage of mathematical concepts, atomic structure, electrochemistry, ionization, equilibrium, acids and bases.

Introduce atomic structure first and build on it. 'If the student will lose interest' is a thought conceived in the idea of the poorer student, and know that the interest was never there under any type of presentation.'

This opinion is countered by Sister M. Lawrence of Pittsburgh in her survey of twenty chemistry teachers east of the Mississippi. These teachers were sampled at the high school and the college level and their attitudes are represented by the following:

"Eight of the twenty authorities consulted believe that the number one trend in teaching chemistry is toward education of the whole person for intelligent citizenship.... They (including Conant of Harvard, French of Colgate, and Taylor of Princeton) feel that science is not being taught and assimilated into our culture to the extent that it should be. They advocate a new approach to secondary teaching to provide time and space for the logical development of scientific thinking. They feel that science courses, consisting of facts, principles, and problems necessary for professional training, allow no place for development of scientific thought or for the understanding of the growth of science. They recommend a course of important historical scientific topics so that the student may study how early problems arose, how they were attacked, and the significance of their solution to society."

According to Lawrence, the course would include the difficulties involved in each new advance in chemistry, the false steps, the new experimental techniques, the relationship of experiment with theory, and the modification and replacement of older theories with new ones.

2/Id., p. 135.
Three others of the scientist-educators polled considered the intangible contributions of chemistry to the development of the character of the student of vastly greater importance than the factual knowledge of chemistry itself.

Nicholson affirms this stand: 1

"The present day trend of stressing the physical approach may be due to the influence of modern developments in nuclear chemistry, the backgrounds and training of the textbook authors, or a combination of both. I firmly believe that descriptive coverage is vital to the introductory course and should not be sacrificed. Could it be that this present trend could be a partial cause of our current scarcity of potential chemistry majors? Are we literally scaring them out of continued study before they have an opportunity to declare their major?"

Thelen, 2 back in the forties, gathered evidence in support of the inductive and theoretical approach, as deductive-descriptive chemical laboratory exercises did not prove the best method for high school instruction. He employed a combination of macro-semimicro studies with deductive-inductive measuring. His control groups were macro and semimicro deductive. The basic underlying principle of the work was, that through employment of instructional activities guided by the experimental method, (inductive teaching) students would learn as much or more subject matter as under the control method, (deductive teaching) and would make substantially greater progress in development of desirable qualities in the area of critical thinking. Since he was dealing with four variables, he treated his data by the analysis of covariance


2/Herbert Thelen, op. cit., p. 68-70.
and found that on a year's study the level of significance in gain among the methods was 1.5 percent in the experimental vs. the semi-micro-deductive group; and 1.0 percent in the experimental vs. the deductive macro.

Boeck has published several articles in the field: 1/

"In the deductive-descriptive method of teaching the experiments were carried out after the general principles involved had been thoroughly discussed in recitation period. No provision was made for student planning of experiments. Little or no opportunity was provided for solution of real problems under laboratory conditions. In the inductive teaching the laboratory was used to obtain data to be used in the solution of a problem the student had a real desire to solve. The problem was used to introduce the general principles of chemistry. The mode of attack was pupil planned with teacher guidance....The inductive did as well or better than the deductive-descriptive in the attainment of the general outcomes of the high school course in chemistry and was significantly superior with respect to the crucial problem of attainment of knowledge and ability to use methods of science with an accompanying scientific attitude."

Boeck further stated that no reference was made by pupils to texts for ideas as to how to proceed during the planning period of an experiment. Consulting references, he felt, would easily provide pupils with the final answer before the evidence was collected, and thus would destroy the chief value of this method, viz., the planning, collection of data and its analysis. He wished his pupils' experiments to become the bases for equation writing, with each pupil contributing to the best of his ability, and all learning from the constant repetition and review provided by the work.

Bayles: 2/ endeavored to refute the "Newtonian Method of Science"

(Philosophic Realism) such as Einstein had in mind when he excoriated the inductive method saying, "There is no inductive method that could lead to the fundamental concepts of physics." In his book, *Out of My Later Years*, Dr. Einstein says:

"Failure to understand this fact constituted the basic philosophical error of so many investigators of the nineteenth century. Physics constitutes a logical system of thought which is in a state of evolution and whose bases cannot be established through distillation by any inductive method from the experience learned through, but which cannot be attained by, free investigation."

Bayles claims Einstein had in mind induction in its strictest sense, viz., the process of gathering facts and letting the facts speak; of reading nothing of personal opinion or bias into them. Bayles himself feels that the truly inductive method of "progressing from the particular to the general, or more specifically from facts to concepts or principles" is the one by which scientific knowledge has grown from its earliest beginnings. Hence, inductive thinking in this restricted sense definitely has its place in chemistry teaching.

The concluding note in this overall survey of related literature in the field of inductive chemistry is sounded by Mallinson and Buck in the *Journal of Chemical Education*. They cite the aspects of laboratory work that have been under consideration and criticism in later years, chief among them being the slight change in program over the years when the field of chemistry itself has changed so greatly.


2/Ernest Bayles, op. cit., p. 115

the woeful na"ivity of high school students in the use of the scientific method; and the sheer mechanical way in which students balance chemical equations without understanding the significance of the reactions involved. They feel that students, as a rule, are unable to think chemically. The emphasis of the typical laboratory manual is all on the deductive-imitative approach which requires the student to draw few if any conclusions. In the last analysis the experiments amount to little more than "cookbook" manipulations, the results of which have to be written in blanks. Wallinson feels the inductive approach should be used in all phases of science teaching and outlines the manner in which he employs it

"Nitrogen dioxide is one of a number of oxides of nitrogen. It's heavier than air and soluble in water. It's prepared by copper and concentrated nitric acid. I'll give you an amount of solution that contains 20 grams of HNO₃. I want you to set up an apparatus in which this material can be prepared and collected safely. I want you to figure out how many grams of copper shot you will need."

The question naturally arises as to whether these experiments will not take much longer than the traditional deductive-descriptive ones. Wallinson and Buck have commented on several disputed topics such as this, and frankly admit that inductive experimentation is more time consuming. But, that is an advantage, in that it teaches the student to think. On the question of the difficulty of an inductive course they remark that such a course is not "automatically tougher". The big advantage they feel, is that the reduction in the number of imitative experiments will give the students more time to do what chemistry

\[\text{George Wallinson and Jacqueline Buck, op. cit., p. 635.}\]
teachers have been paying lip-service to for so long, viz., to solve chemical problems.1/

"Present the problem of the laboratory exercises before the topic is discussed in class. If the chemistry laboratory is designed to introduce the students to the techniques used by chemists, then it must be given prior to classroom discussion. Do chemists only attack problems to which they know the answer?"

Such a system in the chemistry laboratory will necessitate a reduction in the number of imitative exercises in order to allow for exercises in thinking, and this is heartily to be commended. The interest of the student will be aroused because of the recognizable purpose of his efforts. However, problems of importance to the student must be selected, not those considered "essential" by some textbook author. Also, the teacher must demand further accuracy as the course progresses.2/

Nollinson closed his paper with the words:3/

"Perhaps further study of these problems together with efforts to use some of the suggestions here will make the chemistry laboratory more meaningful than heretofore. As emphasized by Schlesinger, "It is essential to keep in mind the major objectives--training in observation, in thought and in action."

1/George Nollinson and Jacqueline Buck, op. cit., p. 587
2/Ibid., p. 632.
CHAPTER III

PHASE I - RESEARCH PROCEDURE AND DATA

Phase I of this study consists in a comparison of inductive and deductive methods of teaching a full year's course in high school chemistry. Data were obtained from twelve classes located in three separate schools, taught by three different teachers.

Chapter III describes the research methods employed in the twelve classes used within these three schools in the period during which the experiment was conducted.

The study was designed to show whether or not any significant statistical difference exists in the knowledge of general chemistry which a student acquires when taught inductively or deductively. The tools used to measure the knowledge acquired were the Anderson¹ and the Cooperative² standardized chemistry tests. On all twelve classes complete data on achievement and mental capacity were collected.

This chapter will set forth a description of the textbooks, laboratory manuals and guide sheets used in the classes, and the syllabus followed by the teachers. The time schedule, the conditions, subjects and procedures used in each of the classes are shown. A summary of the data collected is followed by a display of this data in chart form in order that the reader may form a picture of the study.

¹/Anderson Chemistry Test, op. cit.
²/Cooperative Chemistry Test, op. cit.
1. Description of Procedure and Materials

Textbooks and manuals.-- Test classes A, B, C, and D (deductive chemistry classes) were grade eleven, college preparatory rooms with thirty-five girls in each, at the Monsignor Ryan Memorial High School, in Dorchester, Massachusetts. (This school is herein designated as School A.)

To each girl was given a standard laboratory workbook and laboratory equipment of the type generally supplied in secondary schools. The laboratory classes took place during two consecutive periods per week, and the girls worked individually in laboratory sections of twelve, with the classroom teacher supervising.

The classroom program followed the traditional pattern of lecture and drill on descriptive chemistry and theoretical chemistry. The syllabus used was that of the Archdiocese of Boston Parochial School System, 1946 revision, (see Appendix A) supplemented with and modified by reference to the American Chemical Society Syllabus of 1948 as outlined for the Middle States Secondary Schools, as well as the New England Association of Chemistry Teachers' Syllabus revised in 1943.

There were no demonstrations given in any of the test classes nor were any visual aids employed beyond the regular chalkboard and the Periodic Chart of the elements.

A textbook was supplied to every student and assignments were


made daily from the exercises contained therein. Forty-five minute examinations were given every Friday on the matter covered in the class and laboratory sections during the week; a review examination was given each quarter. Marks were computed from these weekly tests and quarterly examinations with the quarterlies counting double value. In April and May two standardized chemistry tests were given. Intelligence test records yielded IQ scores from the Sophomore year.

Test classes E and F (inductive chemistry) were taught by a different classroom and laboratory method. The same textbook (1934 edition) was used, and in addition, copies of every available high school text were placed on a shelf in the laboratory. The Chemical Rubber Company Handbook and several college texts on general chemistry were also included.

Syllabus.— The syllabus followed was designed by the writer after she had made a survey of the fifteen colleges and nursing schools in the East most commonly accepting girls from the area being tested. Questionnaires were sent to the chemistry teachers of the institutions asking them what topics they felt should be stressed on the secondary level. These recommendations, coupled with the College Entrance Board requirements and suggestions published by the American Chemical Society were all studied and worked over until the syllabus was devised. (Appendix B) (Note: This syllabus was discussed at a meeting of chemistry teachers from the Midwest at Pittsburgh in the summer of 1953.

1/Anderson Chemistry Test, op. cit.
2/Copernicus Chemistry Test, op. cit.
and was later incorporated, in large measure, into the chemistry syllabus for a school district of that area.)

The classes were conducted for five periods per week and a new laboratory approach was tried. The students were given the topic of the week's laboratory work one week in advance and were left to themselves to read up on the subject and discover the procedures to use. There was no restriction of source material; the reference shelf in the laboratory was available for use in the building, and all were advised to make use of the public library facilities in addition.

During the year it became evident that some of the students were merely copying directions from old laboratory manuals and, though it was suggested that they might learn more by other means, no positive restrictive steps were taken. Students were also free to use any equipment that was in the laboratory, whether it was macro or semimicro. Since they found all their references in the macro language, they used that size sample almost entirely.

**Testing.**--In May and June of the school year, standardized chemistry tests were administered. Intelligence test scores from the Sophomore year were recorded. Weekly tests on the matter covered in the classroom and the laboratory were given, and particular stress was laid on the laboratory learning since it was desirable to know how much of the preparations, properties and uses of elements the students had learned for themselves. The classroom lectures covered pure theory only and left the descriptive part of chemistry for the

1/ California Test of Mental Maturity, op. cit.
2/ Anderson and Cooperative Chemistry Tests, op. cit.
student to learn by reading and experimenting. Quarterly tests were also given, and these counted double in making out the report card grades.

**Inductive semimicro classes.** Classes C and H (inductive) were taught by the same syllabus as classes E and F just described, but were allowed only semimicro procedures in laboratory work. Centrifuges were introduced; advances into chromatography were encouraged and guide sheets were provided for experiments to be done. (Appendix C) In addition, a general element-compound outline was given to each student. (See page 9) This outline contained a sketch of the minimum amount of information that had to be known about each element or compound studied and hence made it easier for the student to limit his laboratory work to pertinent information.

**Time schedule.** In September, techniques were demonstrated to all laboratory classes and students familiarized themselves with the handling of chemical apparatus before they were permitted to start experimenting on their own. These basic instructions included weighing, centrifuging, filtering, pouring, collecting gases by displacement of air and of water, and the testing of gases, liquids and solids.

2. Outside Schools

**Test School B.** The first outside school used in the study and herein designated as School B, was a high school in the West Indies. The chemistry classes were taught by an American teacher with a Bachelor's and Master's degree in chemistry. The classes consisted of white children mixed male and female, and of statistically comparable age and intelligence
to the Dorchester school. The teacher had several conferences with the writer in the summer and continued contact by correspondence. The island instructor took back with her copies of the inductive course of study (Appendix B) and laboratory guide sheets (Appendix C) and taught with them as far as possible. The general procedures used were the same as in the Dorchester school and the Cooperative and Anderson chemistry tests were written in April and May. The laboratory had general equipment but much of the laboratory work was done on the semimicro scale since chemicals are scarce and hard to store in the tropics. Data from the previous year were taken as a deductive standard class against which to compare the experimental group. Records of the Anderson and of the Cooperative chemistry tests were also available for this class. The same teacher had taught both classes.

Test School C.-- The second test-school, herein designated as School C, was located in an urban center in Ohio. The teacher in this case had also undergraduate and graduate degrees in chemistry. The students were mixed male and female, and of statistically comparable age and intelligence to those of the Dorchester school. In this school the syllabus and guide sheets of the inductive method were also followed and the teacher kept in close contact with the writer by mail. Slight modifications were introduced into the syllabus because of local American Chemical Society recommendations. The additions were in the field of organic chemistry. The Cooperative and the Anderson chemistry tests were administered at the same time as in the other schools (April and May) but the Ohio school worked on the one hour per week laboratory period rather than with the double period allotted to the classes in
the other two schools, The results obtained on the Anderson and the
Cooperative chemistry tests appear in Table 4. The text used in the
Ohio school differed from that used in the other two schools. 1/ Again
in this school the records from the preceding year were taken in order
to have a local norm for comparison. Intelligence test scores for
both classes were obtainable. 2/

3. Conditions, Subjects, and Procedures

The first phase of the study took place at the Monsignor Ryan
Memorial High School in Dorchester, Massachusetts, during the school
years 1949-1950. Eight classes of chemistry students participated in
the study. The essentials of the regular practice were:

1. One class period per day, or five class periods per week, were
given. Each class lasted forty-five minutes.
2. All students were issued the same textbook.
3. All students were given the same daily assignments, the same
lectures and the same demonstrations.
4. Two periods per week were spent in the laboratory.
5. Laboratory assignments were made one week in advance.
6. Laboratory assignments were carried out individually.
7. Chemistry was given to the third year of high school students,
the junior classical, college preparatory division. Classes averaged
thirty-five students. "Adequate supplies of chemicals and glassware

1/Chemistry and You, Lyons Gamahan, Saint Louis, Missouri, 1953.
2/California Test of Mental Maturity, op. cit.
were available and the classroom teacher was present in the laboratory during all laboratory hours.

8. In the years 1948-1950, the students followed experiments and filled in the answers in the workbook *Discovery Problems in Chemistry.*

9. In the years 1948-1952, the regular course in chemistry as outlined by the Archdiocese of Boston Parochial School System for the Boston area was used. (See Appendix A) This course of study covers the traditional areas of secondary school chemistry as well as the newer areas of atomic energy.

10. In the years 1950-1952 the students did semimicro work using a mimeographed copy of the laboratory manual designed by the writer. (See Appendix B.)

11. In the years 1953-1956, the students carried out laboratory assignments with no laboratory manual, but with guide sheets designed by the writer. (See appendix C) After the laboratory work was completed, each student wrote a laboratory report and handed this report in for grading.

12. In the years 1953-1956, the course of study followed by the teachers in the class lectures was designed by the writer. It consists of nine major topics of theoretical chemistry but does not touch on descriptive chemistry at all; this phase of the subject is learned completely in the laboratory. (Appendix B)

\[1/\text{C.E. Fokert, *Discovery Problems in Chemistry*, op. cit.}\]
4. Data Collected

1. The Anderson Chemistry Test 1/ was administered to each group in April of the school year.

2. The Cooperative Chemistry Test 2/ was written by every student in May of the school year.

3. The California Test of Mental Maturity, advanced form, 3/ was written in the Sophomore year and these scores were used.

   All of these are standardized tests and gave data which were comparable with national norms.

5. Data Displayed

The California Test of Mental Maturity 4/ was administered to all students in the groups and the test scores were used for equating the groups. The mean intelligence quotient and the standard deviation for each class were determined. The results appear in Table 3.

Table 3. Means-Standard Deviations, California Test of Mental Maturity,
School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Year</th>
<th>Mean</th>
<th>Standard Deviation</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1942</td>
<td>107.92</td>
<td>8.92</td>
<td>31</td>
</tr>
<tr>
<td>B</td>
<td>1949</td>
<td>109.05</td>
<td>10.50</td>
<td>34</td>
</tr>
<tr>
<td>C</td>
<td>1950</td>
<td>111.48</td>
<td>3.33</td>
<td>32</td>
</tr>
<tr>
<td>D</td>
<td>1951</td>
<td>113.36</td>
<td>12.20</td>
<td>35n</td>
</tr>
<tr>
<td>E</td>
<td>1952</td>
<td>111.52</td>
<td>2.12</td>
<td>37</td>
</tr>
<tr>
<td>F</td>
<td>1953</td>
<td>111.06</td>
<td>10.64</td>
<td>40</td>
</tr>
<tr>
<td>G</td>
<td>1954</td>
<td>105.24</td>
<td>3.32</td>
<td>40</td>
</tr>
<tr>
<td>H</td>
<td>1955</td>
<td>109.11</td>
<td>9.60</td>
<td>33</td>
</tr>
</tbody>
</table>

1/ Anderson Chemistry Test, op. cit.

2/ Cooperative Chemistry Test, op. cit.

3/ California Test of Mental Maturity, op. cit.

4/ Ibid.
During the third school week of April the Anderson Chemistry Test was administered and scored. The means-standard-deviations were calculated for each class and appear in the following table.

Table 4. Means-Standard-Deviations, Anderson Chemistry Test, School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Year</th>
<th>Mean</th>
<th>S.D.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1948</td>
<td>60.25</td>
<td>22.64</td>
<td>31</td>
</tr>
<tr>
<td>B</td>
<td>1949</td>
<td>56.30</td>
<td>25.20</td>
<td>34</td>
</tr>
<tr>
<td>C</td>
<td>1950</td>
<td>62.20</td>
<td>17.70</td>
<td>32</td>
</tr>
<tr>
<td>D</td>
<td>1951</td>
<td>62.45</td>
<td>18.55</td>
<td>35</td>
</tr>
<tr>
<td>E</td>
<td>1952</td>
<td>70.55</td>
<td>12.05</td>
<td>37</td>
</tr>
<tr>
<td>F</td>
<td>1953</td>
<td>80.10</td>
<td>13.14</td>
<td>40</td>
</tr>
<tr>
<td>G</td>
<td>1954</td>
<td>80.50</td>
<td>14.95</td>
<td>40</td>
</tr>
<tr>
<td>H</td>
<td>1955</td>
<td>74.94</td>
<td>18.30</td>
<td>33</td>
</tr>
</tbody>
</table>

During the second week of May the Cooperative Chemistry Test was given and scored. The results were used to calculate the means and standard deviations which appear in Table 5.

Table 5. Means-Standard-Deviations, Cooperative Chemistry Test, School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Year</th>
<th>Mean</th>
<th>S.D.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1948</td>
<td>70.50</td>
<td>18.20</td>
<td>31</td>
</tr>
<tr>
<td>B</td>
<td>1949</td>
<td>67.25</td>
<td>17.00</td>
<td>34</td>
</tr>
<tr>
<td>C</td>
<td>1950</td>
<td>71.00</td>
<td>14.65</td>
<td>32</td>
</tr>
<tr>
<td>D</td>
<td>1951</td>
<td>67.75</td>
<td>16.00</td>
<td>35</td>
</tr>
<tr>
<td>E</td>
<td>1952</td>
<td>83.90</td>
<td>7.50</td>
<td>37</td>
</tr>
<tr>
<td>F</td>
<td>1953</td>
<td>87.15</td>
<td>11.05</td>
<td>40</td>
</tr>
<tr>
<td>G</td>
<td>1954</td>
<td>84.40</td>
<td>15.27</td>
<td>40</td>
</tr>
<tr>
<td>H</td>
<td>1955</td>
<td>83.90</td>
<td>20.50</td>
<td>33</td>
</tr>
</tbody>
</table>
The other two schools employed in this part of the study observed the same time schedule, administered the same tests, and sent the records to the writer. These scores have been tabulated, and the means and standard deviations have been calculated. Results of the two schools are shown in Table 6.

Table 6. Intelligence and Achievement Test Scores - Schools B and C

<table>
<thead>
<tr>
<th>Class</th>
<th>Intelligence Test Mean</th>
<th>S.D.</th>
<th>Anderson Test Mean</th>
<th>S.D.</th>
<th>Cooperative Test Mean</th>
<th>S.D.</th>
</tr>
</thead>
<tbody>
<tr>
<td>School B</td>
<td>(1)</td>
<td>111.42</td>
<td>10.56</td>
<td>52.47</td>
<td>10.02</td>
<td>54.11</td>
</tr>
<tr>
<td>A .....</td>
<td>(2)</td>
<td>9.42</td>
<td></td>
<td>46.67</td>
<td>5.28</td>
<td>49.61</td>
</tr>
<tr>
<td>B .....</td>
<td>(3)</td>
<td>53.99</td>
<td>10.20</td>
<td>59.35</td>
<td>11.70</td>
<td></td>
</tr>
<tr>
<td>School C</td>
<td>(4)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>A .....</td>
<td>109.91</td>
<td>7.93</td>
<td>46.67</td>
<td>5.28</td>
<td>49.61</td>
<td>4.38</td>
</tr>
<tr>
<td>B .....</td>
<td>109.23</td>
<td>6.96</td>
<td>53.99</td>
<td>10.20</td>
<td>59.35</td>
<td>11.70</td>
</tr>
</tbody>
</table>

High School grades in chemistry and general science, and general averages were also available for counting the groups, but in considering these grades among the various schools, it was discovered that too much subjectivity in marking made the use of such criteria of questionable value.

Experimental and control groups. -- The deductive-descriptive classes were designated as control groups and the inductive classes as experimental groups. Hence, in the Dorchester school, classes A, B, C, and D were controls, whereas classes E, F, G, and H were experimental. Ohio class A and West Indian class A were controls; Ohio class B and West Indian class B were experimental. Both control groups and experimental groups received instruction in the same topics and units in the general high school chemistry course but a different method of
approach was taken in the case of the inductive experimental groups.

The reader has now seen an overall picture of Phase I of this study, both in the verbal descriptions given and in the data displayed in the foregoing charts. Chapter IV will proceed with an analysis of this data.
CHAPTER IV
ANALYSIS OF DATA

Statistical tools had to be chosen for the analysis of the data collected by the administration of tests during the study. A satisfactory tool for this work must possess the following characteristics:

1. It must be able to test the significance of difference between two groups having slightly different numbers of individuals.

2. It must be capable of adjusting the data to the criteria means if the control variations actually affect the criteria.

3. It must be sensitive enough to test the difference between means if true differences actually exist.

To test the null hypothesis that no true differences exist between groups in terms of general ability, critical ratios were determined among the scores on the California Test of Mental Maturity.

Analysis of the scores on the California Test of Mental Maturity.

The scores of the test were tabulated and critical ratios were determined among the class medians of each school, inductive and deductive, and among the total inductive and total deductive classes in all three schools. These results appear in the following tables. Table 7 lists the results in test School A while Table 8 contains the figures obtained in test schools B and C.

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The tables were constructed from the data received from the twelve classes in the three schools of the study. Since classes were approximately equal in size, the weighting of scores to obtain the combined means is not as apparent in class charts as in those of total school results. The formula used for the calculation of the combined means:

\[ M_{\text{comb.}} = \frac{N_1 M_1 + N_2 M_2 + \ldots + N_m M_m}{N_1 + N_2 + \ldots + N_m} \]

In which \( N_1, N_2, \text{etc.} \) represent the numbers of pupils in the classes, and \( M_1, M_2, \text{etc.} \) are the means of the class scores.

For the combined standard deviations the following formula was used:

\[ SD_{\text{comb.}} = \sqrt{\frac{N_1 (SD_1^2 + d_1^2) + N_2 (SD_2^2 + d_2^2) + \ldots + N_m (SD_m^2 + d_m^2)}{N_1 + N_2 + \ldots + N_m}} \]

\( N_1, N_2, \text{etc.} \) represent the number in a class or school
\( M_1, M_2, \text{etc.} \) represent the mean scores
\( SD_1, SD_2, \text{etc.} \) represent the standard deviations of class or school
\( d_1, d_2, \text{etc.} \) represent the difference between the means of the group and the combined mean

**Application.**—In Table 7, the application of the formula to the data yields the following combined mean score:

\[ M_{\text{comb.}} = \frac{31(107.92) + 34(109.05) + 32(111.48) + 35(113.32)}{31 + 34 + 32 + 35} = 110.51 \]


2/ Ibid. p. 57.
The term "difference" as used in the following tables signifies the arithmetical difference between the two mean scores under comparison, e.g. In Table 7, the difference = 110.51 - 109.33 = 1.18.

In order to discover whether the groups differ sufficiently in mean performance to enable us to say with confidence that the difference will persist upon repetition of the experiment, the standard error between the two means must be known. This error is calculated from the formula:

\[ SE_D = \sqrt{\frac{\sigma_1^2}{N_1} + \frac{\sigma_2^2}{N_2}} \]

When \( \sigma_1 \) and \( \sigma_2 \) are the standard deviations of the groups, and \( N_1 \) and \( N_2 \) are the numbers of pupils in the groups.

Applied to the data of Table 7:

\[ SE_D = \sqrt{\frac{0.48^2}{132} + \frac{0.52^2}{150}} \]

\[ SE_D = 1.10 \]

The critical ratio is obtained by dividing the difference by the standard error, as \( CR = \frac{D}{SE_D} \). In this case the \( CR = 1.18/1.10 \) which is 1.07.

When a critical ratio is 1.96 or higher, a null hypothesis is rejected at the five percent level of confidence; if the critical ratio is 2.58 or more, it is rejected at the one percent confidence level.

The t-test used in Table 10 and the tables beyond is a general test of significance used with small groups. The t-value is calculated from the formula:

\[ t = \frac{D}{SE_D} \]

1/ Henry E. Garrett, op. cit. p 214.
2/ Ibid. p. 217.
Table 7. Critical Ratios of Mental Maturity Tests - School A

| Class | Deductive | | | | | Inductive | | | | |
|-------|-----------|---|---|---|---|---|---|---|---|---|---|
|       | Mean      | S.D. | S.E. | N  |   | Mean      | S.D. | S.E. | N  |   |
| (1)   | (2)       | (3) | (4) | (5) |   | (6)       | (7) | (8) | (9) |   |
| A     | 107.92    | 8.82| 1.00| 31 |   | E         | 111.52| 9.12| 1.50| 37 |   |
| B     | 109.07    | 10.50| 1.80| 34 |   | F         | 111.06| 10.64| 1.70| 40 |   |
| C     | 111.48    | 9.33 | 0.3  | 32 |   | G         | 105.24| 8.32 | 1.30| 40 |   |
| D     | 113.36    | 12.20| 2.1  | 35 |   | H         | 109.11| 9.60 | 1.70| 33 |   |
| Average| 110.51   | 9.40 | 1.20 |   |   | 109.53   | 9.50 | 1.55 |   |   |

| Difference * | 1.18 |
| S.E. D ** | 1.10 |
| C.R. = D/SE D | 1.07 |

Hence, we have no statistical difference in IQ scores at the one or five percent levels of confidence, and we may assume these groups to be considered as members of the same population.

In a similar fashion the critical ratios were determined for the outside schools B and C. The results appear in Table 8.

Table 8. Critical Ratios of Mental Maturity Tests - Schools B and C

| Class | Deductive | | | | | | | | | | |
|-------|-----------|---|---|---|---|---|---|---|---|---|---|---|
| School B | (1) | (2) | (3) | (4) |   | School C | | | | |
| A     | 111.42 | 10.56 | 1.81 | 35 |   | A     | 110.91 | 7.85 | 1.42 | 31 |
| A     | 110.22 | 9.42 | 1.62 | 35 |   | B     | 109.28 | 6.96 | 1.22 | 33 |

| Difference (School B) | 1.20 |
| S.E. D ** | 2.43 |
| C.R. = D/SE D | 0.49 |

| Difference (School C) | 1.63 |
| S.E. D ** | 1.87 |
| C.R. = D/SE D | 0.87 |

* Difference between means.
** Standard error of the difference between the two means.
Again there is no significant variation in the intelligence test scores at the 0.05 or 0.01 levels of significance, so we shall now group all inductive classes in the three schools and all deductive classes in the three schools, and determine the critical ratio of the composite intelligence test scores.

Table 9. Determination of the C.R. for Inductive and Deductive Classes in the Three Schools Involved in the Study (Intelligence Tests)

<table>
<thead>
<tr>
<th>School</th>
<th>Inductive Mean IQ</th>
<th>S.E.</th>
<th>School</th>
<th>Inductive Mean IQ</th>
<th>S.E.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(1)</td>
<td>(2)</td>
<td></td>
<td>(4)</td>
<td>(5)</td>
</tr>
<tr>
<td>A</td>
<td>110.43</td>
<td>1.75</td>
<td>A</td>
<td>109.23</td>
<td>1.55</td>
</tr>
<tr>
<td>B</td>
<td>111.42</td>
<td>1.81</td>
<td>B</td>
<td>110.72</td>
<td>1.62</td>
</tr>
<tr>
<td>C</td>
<td>110.91</td>
<td>1.42</td>
<td>C</td>
<td>109.20</td>
<td>1.22</td>
</tr>
<tr>
<td>Average</td>
<td>110.73</td>
<td>1.53</td>
<td></td>
<td>109.73</td>
<td>1.46</td>
</tr>
</tbody>
</table>

Difference: 1.15       N (deductive): 198

\( \delta \gamma \): 2.11       N (inductive): 218

C.R.: 0.54

The scores obtained on the Anderson Chemistry Test were then compared as to variation in scores with method of teaching. Since the critical ratios of the IQ scores showed the classes to be of comparable ability, variations in the Anderson Test scores may be investigated. The t-test was applied to test the null hypothesis that inductive laboratory learning produces no different achievement on the Anderson Chemistry Test than does deductive-descriptive learning. Since the numbers involved are not large, the t-test was considered to be the best statistical tool. The analysis of data for
test school A appears in Table 10.

Table 10. Analysis of Data on the Anderson Chemistry Test - School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>No.</th>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td></td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
</tr>
<tr>
<td>A</td>
<td>60.25</td>
<td>22.05</td>
<td>4.5</td>
<td>31</td>
<td>F</td>
<td>70.55</td>
<td>12.65</td>
<td>2.1</td>
<td>37</td>
</tr>
<tr>
<td>B</td>
<td>58.50</td>
<td>25.50</td>
<td>4.7</td>
<td>34</td>
<td>G</td>
<td>60.10</td>
<td>15.14</td>
<td>2.1</td>
<td>40</td>
</tr>
<tr>
<td>C</td>
<td>62.50</td>
<td>17.70</td>
<td>3.2</td>
<td>37</td>
<td>H</td>
<td>74.95</td>
<td>18.30</td>
<td>3.1</td>
<td>43</td>
</tr>
<tr>
<td>D</td>
<td>62.45</td>
<td>13.55</td>
<td>3.0</td>
<td>35</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Average</td>
<td>60.45</td>
<td>21.03</td>
<td>5.6</td>
<td>33</td>
<td></td>
<td>70.52</td>
<td>14.61</td>
<td>2.4</td>
<td>33</td>
</tr>
</tbody>
</table>

S.E. = 17.90
S.E. = 4.12

\[ t = \frac{D}{S.E.} = 4.39 \]

Degrees of freedom = 69

From tables: \( t = 2.00 \) at the 0.05 level; \( t = 2.65 \) at the 0.01 level.

A \( t \)-value of 4.39 with 69 degrees of freedom is significant at the one percent level of confidence; therefore, the null hypothesis is rejected and apparently true differences are associated with the use of different methods of teaching chemistry. The difference in favor of the inductive method seems to indicate that this method is superior in terms of general chemical knowledge acquired, as measured by the Anderson Chemistry Test.

The data from test school B, located in the West Indies, were similarly treated and the results appear in Table 11 on the page following.
A t-value with 68 degrees of freedom must equal at least 2.00 to be significant at the five percent level; also, it must equal or exceed 2.66 to have significance at the one percent level. The t-score for this test measures 2.26 which has significance at the five percent level of confidence. Therefore, we reject the null hypothesis on the grounds that not more than once in twenty trials would a difference as large or larger than this occur if the true difference were zero.

Table 11. Analysis of Data on the Anderson Chemistry Test – School B

<table>
<thead>
<tr>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.E.</th>
<th>S.E.</th>
<th>No.</th>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.E.</th>
<th>S.E.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
<td>(9)</td>
<td>(10)</td>
</tr>
<tr>
<td>A</td>
<td>54.27</td>
<td>10.11</td>
<td>1.73</td>
<td>35</td>
<td>B</td>
<td>60.20</td>
<td>11.64</td>
<td>2.00</td>
<td>35</td>
</tr>
</tbody>
</table>

Difference ——— 5.95
S.D. ———— 10.90
S.E. ———— 2.62

Degree of freedom — 68

From tables: t = 2.00 at the 0.05 level; t = 2.66 at the 0.01 level.

Table 12. Analysis of Data on the Anderson Chemistry Test – School C

<table>
<thead>
<tr>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.E.</th>
<th>S.E.</th>
<th>No.</th>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.E.</th>
<th>S.E.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
<td>(9)</td>
<td>(10)</td>
</tr>
<tr>
<td>A</td>
<td>46.87</td>
<td>5.20</td>
<td>0.96</td>
<td>31</td>
<td>B</td>
<td>55.09</td>
<td>10.20</td>
<td>1.87</td>
<td>35</td>
</tr>
</tbody>
</table>

Difference ——— 6.42
S.D. ———— 2.19
S.E. ———— 2.05

Degree of freedom — 62

From tables: t = 2.00 at the 0.05 level; t = 2.66 at the 0.01 level.
Hence, a t-value of 3.13 is significant at the one percent level of confidence. We are justified in concluding that the superiority of the inductive method is upheld in this school as measured by the results on the Anderson Chemistry Test. Again we reject the null hypothesis. The inductive teaching does not bring about significantly different results than deductive teaching.

A composite was then made of all the scores of the inductive classes and all the scores of the deductive classes. That test was applied to these scores on the Anderson Chemistry Test. The work is summarized in Table 13.

Table 13. The t-test Applied to Scores on the Anderson Chemistry Test
Schools A + B + C

<table>
<thead>
<tr>
<th>School</th>
<th>Deductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>No.</th>
<th>School</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
<td>(9)</td>
</tr>
<tr>
<td>A</td>
<td>60.43</td>
<td>21.03</td>
<td>3.60</td>
<td>33</td>
<td>A</td>
<td>78.52</td>
<td>14.61</td>
<td>2.57</td>
<td>38</td>
</tr>
<tr>
<td>B</td>
<td>54.27</td>
<td>10.11</td>
<td>1.73</td>
<td>35</td>
<td>B</td>
<td>60.20</td>
<td>11.64</td>
<td>2.00</td>
<td>35</td>
</tr>
<tr>
<td>C</td>
<td>46.67</td>
<td>5.28</td>
<td>.96</td>
<td>31</td>
<td>C</td>
<td>53.09</td>
<td>10.20</td>
<td>1.87</td>
<td>33</td>
</tr>
<tr>
<td>Average</td>
<td>56.88</td>
<td>16.62</td>
<td>2.85</td>
<td></td>
<td>Average</td>
<td>71.23</td>
<td>13.30</td>
<td>2.22</td>
<td></td>
</tr>
<tr>
<td>S.D.</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>S.D.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>S.E.</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>S.E.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Difference</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Difference</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>t</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>t</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

A t-value of 7.03 is significant at the one percent level of confidence and we are justified in concluding that the superiority of the inductive method is upheld in the combined schools as measured by the results on the Anderson Chemistry Test. Again we reject the
of the inductive method is upheld in the combined schools as measured by results on the Anderson Chemistry Test. Again we reject the null hypothesis that inductive teaching does not bring about significantly different results than deductive teaching.

The second test used to measure achievement in both groups, inductive and deductive, was the Cooperative Chemistry Test. Form X. This test was administered to all classes in May of each year, scored, and the t-test applied. The results of Test School A appear in Table 14.

Table 14. The t-test Applied to the Cooperative Chemistry Test-School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Deductive</th>
<th>Inductive</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mean</td>
<td>SD</td>
</tr>
<tr>
<td>-------</td>
<td>-----------</td>
<td>-----------</td>
</tr>
<tr>
<td>1</td>
<td>70.50</td>
<td>19.26</td>
</tr>
<tr>
<td>2</td>
<td>67.75</td>
<td>17.00</td>
</tr>
<tr>
<td>3</td>
<td>74.00</td>
<td>14.65</td>
</tr>
<tr>
<td>4</td>
<td>67.75</td>
<td>16.66</td>
</tr>
<tr>
<td>Average</td>
<td>69.25</td>
<td>16.46</td>
</tr>
</tbody>
</table>

n = 14.25
SD = 11.21
SE = 2.58
t = 5.52
Degrees of freedom = 69

From the table: t = 2.00 at the .05 level; t = 2.65 at the .01 level.

A t-value of 5.52 is significant at the one percent level of confidence. We are justified in concluding that the superiority of the inductive method is upheld in this school as measured by the results on the Cooperative Chemistry Test. Again we reject the null hypothesis.
that inductive teaching does not bring about significantly different results than deductive teaching.

Similar treatment of data from the Cooperative test administered in May in the West Indian school (School B) failed to corroborate the above evidence as the t-value was too small to be of value at the five percent level. The results may be seen in Table 15.

Table 15. The t-test Applied to Scores on the Cooperative Chemistry Test
School B

<table>
<thead>
<tr>
<th>Class</th>
<th>Deductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>N.</th>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>S.E.</th>
<th>N.</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
<td>(9)</td>
<td>(10)</td>
</tr>
<tr>
<td>A</td>
<td>67.11</td>
<td>10.11</td>
<td>1.75</td>
<td>35</td>
<td>B</td>
<td>67.13</td>
<td>9.90</td>
<td>1.70</td>
<td>35</td>
</tr>
</tbody>
</table>

Difference ------- 3.02
S.D. ----------- 10.05
S.E. ----------- 2.41

From Tables: t = 2.90 at the 0.05 level; t = 2.15 at the 0.01 level

The t-value here is too small to be significant at either level, hence, in this case, as measured by the Cooperative Chemistry Test, the null hypothesis is sustained and the descriptive-deductive teaching appears in no way different from inductive teaching, in results.

Test school C also administered the Cooperative Chemistry Test, from X, in May and the results of this were forwarded to the writer. The t-test was applied; the data resulting from this application are listed in Table 16.
Table 16. The t-test Applied to Scores on the Cooperative Chemistry Test - School C

<table>
<thead>
<tr>
<th>Class</th>
<th>Deductive</th>
<th>Inductive</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mean</td>
<td>S.D.</td>
<td>S.E.</td>
</tr>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
</tr>
<tr>
<td>A</td>
<td>49.61</td>
<td>4.30</td>
</tr>
<tr>
<td>S.D.</td>
<td>8.72</td>
<td></td>
</tr>
<tr>
<td>D</td>
<td>9.74</td>
<td></td>
</tr>
<tr>
<td>S.E.</td>
<td>2.18</td>
<td></td>
</tr>
<tr>
<td>t</td>
<td>4.46</td>
<td></td>
</tr>
<tr>
<td>Degrees of freedom</td>
<td>62</td>
<td></td>
</tr>
</tbody>
</table>

Again the t-value is significant at the one percent level of confidence. It can be asserted that the null hypothesis of equivalence of inductive and deductive methods must again be rejected. The superiority of the inductive method is again apparent.

Next, data on inductive method classes from the same three schools were compared with those on deductive method classes for the three schools and the t-test was again used to measure the validity of the null hypothesis in regard to inductive and deductive methods of teaching chemistry.

Table 17. The t-test Applied to Scores on the Cooperative Chemistry Test - Schools A+B+C

<table>
<thead>
<tr>
<th>School</th>
<th>Deductive</th>
<th>Inductive</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mean</td>
<td>S.D.</td>
<td>S.E.</td>
</tr>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
</tr>
<tr>
<td>A</td>
<td>69.25</td>
<td>16.49</td>
</tr>
<tr>
<td>B</td>
<td>64.11</td>
<td>10.81</td>
</tr>
<tr>
<td>C</td>
<td>49.61</td>
<td>4.30</td>
</tr>
<tr>
<td>Average</td>
<td>65.12</td>
<td>13.52</td>
</tr>
</tbody>
</table>
From the foregoing data the standard deviation was calculated and gave the value 13.43. The $s.E._D$ is 2.13 and the actual difference between the means 11.65. Using the general formula for $t = \frac{D}{s.E._D}$, the $t$-value obtained was 5.34, which, with 67 degrees of freedom is significant at the one percent level of confidence.

The $t$-value obtained from a comparison of all inductive and deductive classes rejects the null hypothesis that no difference in achievement on the Cooperative Chemistry Test is noticeable under deductive and inductive teaching. The inductive method apparently does produce statistically significantly higher results as measured by this test.

In order to determine the effect of teaching method on the higher and lower intelligences, comparisons were made of inductive and deductive scores in the upper and in the lower quartile as measured by an intelligence test. The results appear in the following tables.

Table 15. The $t$-test, Lower Quartile IQ, Anderson Chemistry Test

<table>
<thead>
<tr>
<th>Class</th>
<th>Deductive</th>
<th>Inductive</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mean (2)</td>
<td>S.D. (3)</td>
</tr>
<tr>
<td>A+B+C</td>
<td>47.55</td>
<td>2.60</td>
</tr>
<tr>
<td>D</td>
<td>17.38</td>
<td></td>
</tr>
<tr>
<td>S.D.</td>
<td>19.02</td>
<td></td>
</tr>
<tr>
<td>$s.E._D$</td>
<td>4.18</td>
<td></td>
</tr>
<tr>
<td>$t$</td>
<td>4.13</td>
<td></td>
</tr>
</tbody>
</table>

1/ California Test of Mental Maturity, op. cit.
A t-value with 74 degrees of freedom is significant at the one percent level of confidence if it exceeds the value 2.65. The lower quartile value of 4.13 is, therefore, significant at the one percent level of confidence. The null hypothesis is rejected and apparently true differences are associated with the use of different methods of teaching with the lower intelligence group. The breakoff point in IQ was a score of 100; the lower limit in the classes here under consideration was 87. The difference in favor of the inductive method suggests that this method is successful with the less intelligent students.

Table 19. Upper Quartile IQ - School A - Anderson Chemistry Test

<table>
<thead>
<tr>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>N</th>
<th>Class</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>N</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>70.00</td>
<td>14.00</td>
<td>55</td>
<td>E</td>
<td>65.64</td>
<td>8.75</td>
<td>55</td>
</tr>
</tbody>
</table>

A composite was made of the lower quartile students from all three schools and a t-value was calculated from these data. The upper quartile t-value of 5.48 is significant at the one percent level of confidence and again the null hypothesis is rejected and the inductive method is found superior in the teaching of the general principles of chemistry as measured by the success on the Anderson Chemistry test.
The upper quartile was then paralleled by a lower quartile composite score and the results are summarized in Table 20.

Table 20. Lower Quartile IQ - Anderson Chemistry Test - Schools A + B + C

<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>A ......</td>
<td>47.35</td>
<td>20.63</td>
<td>38</td>
<td>A ......</td>
<td>64.75</td>
<td>17.60</td>
<td>38</td>
</tr>
<tr>
<td>B ......</td>
<td>53.25</td>
<td>10.49</td>
<td>8</td>
<td>B ......</td>
<td>40.38</td>
<td>3.66</td>
<td>8</td>
</tr>
<tr>
<td>C ......</td>
<td>41.13</td>
<td>4.33</td>
<td>8</td>
<td>C ......</td>
<td>51.13</td>
<td>6.40</td>
<td>8</td>
</tr>
<tr>
<td>Average</td>
<td>45.16</td>
<td>16.90</td>
<td></td>
<td></td>
<td>59.32</td>
<td>13.58</td>
<td></td>
</tr>
</tbody>
</table>

Difference -------- 13.16
S.D. -------------- 15.33
S.E. -------------- 2.51

The t-value thus obtained by using a combined mean and a combined standard deviation was significant beyond the one percent level of confidence thus affirming the statistically significant superiority of the inductive method with the lower quartile of students as chosen from intelligence test scores.

Table 21. Upper Quartile IQ - Anderson Chemistry Test - Schools A + B + C

<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>A ......</td>
<td>70.00</td>
<td>14.00</td>
<td>38</td>
<td>A ......</td>
<td>63.64</td>
<td>8.40</td>
<td>38</td>
</tr>
<tr>
<td>B ......</td>
<td>66.55</td>
<td>2.40</td>
<td>8</td>
<td>B ......</td>
<td>72.58</td>
<td>4.81</td>
<td>8</td>
</tr>
<tr>
<td>C ......</td>
<td>49.50</td>
<td>4.71</td>
<td>8</td>
<td>C ......</td>
<td>63.37</td>
<td>7.52</td>
<td>8</td>
</tr>
<tr>
<td>Average</td>
<td>55.16</td>
<td>11.01</td>
<td></td>
<td></td>
<td>72.14</td>
<td>7.76</td>
<td></td>
</tr>
</tbody>
</table>
From the foregoing data the standard deviation was calculated and found to be 9.51. The standard error was 1.81. With a difference in means of 12.96, this gives a t-value of 12.96 / 1.81 which is 7.17.

A t-value of 7.17 is significant at the one percent level of confidence and we are justified in concluding that the superiority of the inductive method is upheld in the upper quartile intelligence groups, as measured by the California Test of Mental Maturity and the Anderson Chemistry Test. 1/

The data resulting from the administration of the Cooperative Chemistry Test were treated in a similar manner and corroborated the evidence cited in the case of the Anderson Test. 2/ These figures appear in the following tables.

Table 22. The t-test, Lower Quartile IQ, Cooperative Chemistry Test Scores in School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Deductive</th>
<th></th>
<th></th>
<th></th>
<th>Inductive</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
</tr>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
<td>(8)</td>
</tr>
<tr>
<td>A+B+C+D</td>
<td>53.70</td>
<td>16.10</td>
<td>38</td>
<td>E+F+G+H</td>
<td>78.81</td>
<td>14.15</td>
<td>38</td>
</tr>
</tbody>
</table>

Difference ---------------------- 25. 10
S.D. -------------------------- 15.09
S.E.D ------------------------- 3.31
\( t \) ------------------------- 7.57

1/ California Test of Mental Maturity, op. cit.
Anderson Chemistry Test, op. cit.
2/ Ibid.
Table 23. The t-test, Upper Quartile IQ, Cooperative Chemistry Test  
School A

<table>
<thead>
<tr>
<th>Class</th>
<th>Inductive</th>
<th></th>
<th></th>
<th></th>
<th>Class</th>
<th>Inductive</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
<td></td>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
</tr>
<tr>
<td>1</td>
<td>74.05</td>
<td>15.03</td>
<td>33</td>
<td></td>
<td>1</td>
<td>81.40</td>
<td>3.31</td>
<td>33</td>
</tr>
</tbody>
</table>

Difference ------------ 16.6  
S.D. ------------ 10.4  
S.D. ------------ 2.5  
t ------------ 7.1

The t-value of 7.1 is significant at the one percent level of confidence so that the upper quartile intelligence group also indicates the superiority of the Inductive method.

Table 24. The t-test, Lower Quartile IQ, Cooperative Chemistry Test  
Schools A+B+C

<table>
<thead>
<tr>
<th>School</th>
<th>Inductive</th>
<th></th>
<th></th>
<th></th>
<th>School</th>
<th>Inductive</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
<td></td>
<td></td>
<td>Mean</td>
<td>S.D.</td>
<td>No.</td>
</tr>
<tr>
<td>A</td>
<td>53.70</td>
<td>16.10</td>
<td>33</td>
<td></td>
<td>A</td>
<td>70.31</td>
<td>14.15</td>
<td>33</td>
</tr>
<tr>
<td>B</td>
<td>52.25</td>
<td>8.34</td>
<td>8</td>
<td></td>
<td>B</td>
<td>55.50</td>
<td>4.8</td>
<td>8</td>
</tr>
<tr>
<td>C</td>
<td>47.75</td>
<td>3.54</td>
<td>8</td>
<td></td>
<td>C</td>
<td>55.55</td>
<td>9.3</td>
<td>8</td>
</tr>
<tr>
<td>Average</td>
<td>52.64</td>
<td>14.13</td>
<td></td>
<td></td>
<td></td>
<td>72.67</td>
<td>15.49</td>
<td></td>
</tr>
</tbody>
</table>

Difference ------------ 19.23  
S.D. ------------ 15.75  
S.D. ------------ 2.61  
t ------------ 7.36

The t-value of 7.36 is significant at the one percent level of confidence so that again the null hypothesis is rejected and the inductive method is apparently superior as measured by the Cooperative
Chemistry Test on the lower quartile intelligence group of the three schools under consideration in this study.

Table 25. The t-test, Upper Quartile IQ, Cooperative Chemistry Test
Schools A+B+C

<table>
<thead>
<tr>
<th>School</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>No.</th>
<th>School</th>
<th>Inductive Mean</th>
<th>S.D.</th>
<th>No.</th>
</tr>
</thead>
<tbody>
<tr>
<td>A ......</td>
<td>84.80</td>
<td>14.05</td>
<td>30</td>
<td>A ......</td>
<td>91.40</td>
<td>5.51</td>
<td>30</td>
</tr>
<tr>
<td>B ......</td>
<td>69.38</td>
<td>6.32</td>
<td>8</td>
<td>B ......</td>
<td>79.50</td>
<td>7.30</td>
<td>8</td>
</tr>
<tr>
<td>C ......</td>
<td>53.00</td>
<td>5.43</td>
<td>8</td>
<td>C ......</td>
<td>65.15</td>
<td>6.85</td>
<td>8</td>
</tr>
<tr>
<td>Average</td>
<td>76.09</td>
<td>6.27</td>
<td></td>
<td></td>
<td>77.87</td>
<td>10.28</td>
<td></td>
</tr>
</tbody>
</table>

Difference ------------------ 8.21
S.D. ----------------------- 8.52
S.E. ----------------------- 1.62
t -------------------------- 5.06

The results of the t-tests indicate that the inductive method is statistically superior to the deductive-descriptive method with average classes, with students in the lowest quartile in intelligence, and with students in the upper quartile. The data obtained from scores on the Anderson and on the Cooperative chemistry tests agree in revealing this superiority.
CHAPTER V
SUMMARY, CONCLUSIONS, AND IMPLICATIONS

REVIEW:

1. A study of the comparison of inductive and deductive chemistry teaching at the secondary level was made during the years 1948-1956.

2. This phase of the study involved twelve high school classes, eight were classes of thirty-five girls each, in one urban high school in Massachusetts. Two classes of thirty-five students each were located in an urban center in Ohio; and two more classes were in a city in the West Indies.

3. Four hundred and thirty chemistry students in the third year of high school (Grade XI) were used in this study.

4. Half the classes (six) were designated as control groups and were taught by the traditional deductive method; the other half (six) were designated as experimental groups and were taught by the inductive method.

5. The same teacher taught both the inductive and the deductive classes in the same school, thus limiting the variations in teacher motivation to a minimum.

6. All teachers followed the same outline for class presentation of the topics. All administered the same tests.

7. All students following the inductive procedure used the same laboratory guide sheets designed by the writer.
6. All students following the traditional deductive method used the same semimicro laboratory manual sheets designed by the writer, if using semimicro technique, and the same manual and workbook, if using macro technique.

7. The California Test of Mental Maturity was administered to all the students involved in the study.

10. The Anderson Chemistry Test was administered to each class in April of the school year.

11. The Cooperative Chemistry Test was administered to each class in May of the school year.

12. Critical ratios applied among the means-standard deviations of the intelligence test scores among the classes failed to reveal any significant variation in mental ability, hence, variation in chemistry achievement was possibly the result of variation in method of teaching.

13. The t-test was applied among the means-standard deviations of scores on the Anderson Chemistry Test in each school, and then among the total inductive and deductive scores of all three schools. A significant difference in favor of the inductive method was found.

14. The t-test was applied among the means-standard deviations

2/ California Test of Mental Maturity, op. cit.
3/ Anderson Chemistry Test, op. cit.
4/ Cooperative Chemistry Test, op. cit.
of scores on the Cooperative Chemistry Test, Form I., in each school, and then among the three schools, inductive against deductive classes. Significant difference in favor of the inductive approach was found to exist at the one percent level of confidence.

25. In no instance was the deductive approach found to be superior to the inductive approach even at the five percent level of confidence.

Conclusions.--

1. Inductive laboratory teaching produces higher achievement at the one percent level of confidence on the Anderson and Cooperative chemistry tests than does deductive teaching.

2. Students taught inductively apparently are able to grasp fundamental chemical concepts and factual knowledge better than students being taught by the traditional methods.

3. After a year's course in general chemistry, inductively taught, students have a more thorough knowledge of chemistry as measured by the Anderson and Cooperative chemistry tests than do deductively taught students.

Implications.--

1. The further advantages of the inductive approach in terms of increased chemical knowledge warrant investigation.

2. An inductive course of studies designed by one person can with success be followed by others. The two other teachers involved in this study were in close contact with the writer. Further study should be undertaken to determine if total strangers could satisfac-
torily follow the written suggestions and directions for the inductive approach.

3. An inductive approach is more demanding of the resources of the teacher. The amount of chemical background required of a teacher in order to be successful in this field merits more research.

4. Tests designed to measure chemical achievement at the secondary level are few. More investigation should be done in the field of testing understanding of chemical concepts.

5. The writer feels that further studies should be made in the realm of comparison of inductive and deductive methods in relation to extrinsic measures such as correct habits of observation, ability to reason, and the acquisition of scientific attitudes and interests.
CHAPTER VI

PHASE II - RESEARCH PROCEDURE

The second phase of this study is a tightly controlled comparison of inductive and deductive teaching methods on one unit of secondary school chemistry — the concept of equation-balancing.

Scope. — Thirty-one schools scattered across the country were selected. To each of these schools was sent a complete set of notes for the teaching of the unit on chemical equation-balancing by the inductive method. Complete deductive notes were also compiled by the writer and were sent to these teachers also. In these schools, each teacher who taught more than one chemistry class was asked to teach one class inductively and another deductively. In schools where only one class in chemistry existed, teachers were asked to select another teacher of their acquaintance in another school to teach the unit deductively. This meant that, in all, fifty classes were involved in this phase of the study. The classes contained both boys and girls, averaged twenty-five to thirty pupils each, and were composed of third and fourth year of high school students.

Testing. — The students' understanding of the unit was measured by a Symbolic Unit-Test of chemical reactions, devised by the writer. This test was administered before and after the unit was taught. Intelligence test scores for each pupil who participated in the study were also procured.
Table 26. Teachers and Schools Involved in Phase II of the Study

<table>
<thead>
<tr>
<th>School</th>
<th>Locale</th>
<th>Teacher</th>
<th>Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lincoln School</td>
<td>Providence, R.I.</td>
<td>Dorothy Gifford</td>
<td>B</td>
</tr>
<tr>
<td>St. Joseph High</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Innocent, OSF</td>
<td>B</td>
</tr>
<tr>
<td>Foxbury Memorial</td>
<td>Foxbury, Mass.</td>
<td>Joseph Walsh</td>
<td>A</td>
</tr>
<tr>
<td>Bluff City High</td>
<td>Bluff City, Tenn.</td>
<td>Parks Simerley</td>
<td>A</td>
</tr>
<tr>
<td>St. Paul Cathedral</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Annina, RSM</td>
<td>A</td>
</tr>
<tr>
<td>Cathedral High</td>
<td>Canton, Ohio</td>
<td>Sr. Ellen Francis</td>
<td>A</td>
</tr>
<tr>
<td>St. Martin High</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Ida Mary, OSF</td>
<td>A</td>
</tr>
<tr>
<td>Cathedral High</td>
<td>Raleigh, N.C.</td>
<td>Sr. Thomas Mary, OP</td>
<td>A</td>
</tr>
<tr>
<td>Gainesville High</td>
<td>Gainesville, Fla.</td>
<td>Lou Kurtz</td>
<td>A</td>
</tr>
<tr>
<td>Mt. St. Mary Academy</td>
<td>Neshamn, N.Y.</td>
<td>Sr. Gertrude, OSF</td>
<td>B</td>
</tr>
<tr>
<td>Immaculate Conception</td>
<td>Washington, Pa.</td>
<td>Sr. Aloysius, RSM</td>
<td>B</td>
</tr>
<tr>
<td>Pindge Technical</td>
<td>Cambridge, Mass.</td>
<td>John Terry</td>
<td>A</td>
</tr>
<tr>
<td>S.S. Xavier Academy</td>
<td>Latrobe, Pa.</td>
<td>Sr. Theophane, RSM</td>
<td>B</td>
</tr>
<tr>
<td>St. Charles High</td>
<td>New Orleans, La.</td>
<td>Sr. Thomas, OSF</td>
<td>B</td>
</tr>
<tr>
<td>Pius XII High</td>
<td>Passaic, N.J.</td>
<td>Sr. Paul, OP</td>
<td>A</td>
</tr>
<tr>
<td>Lynn View High</td>
<td>Kingsport, Tenn.</td>
<td>John Shearer</td>
<td>A</td>
</tr>
<tr>
<td>Mt. Mercy Academy</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Julia, OSF</td>
<td>B</td>
</tr>
<tr>
<td>Ketron High</td>
<td>Kingsport, Tenn.</td>
<td>Arnold Gilbert</td>
<td>A</td>
</tr>
<tr>
<td>Sweetwater High</td>
<td>National City, Cal.</td>
<td>Charles Dickerhoff</td>
<td>A</td>
</tr>
<tr>
<td>St. John High</td>
<td>Concord, N.H.</td>
<td>Sr. Bridge, OSF</td>
<td>A</td>
</tr>
<tr>
<td>Forest Hills High</td>
<td>Forest Hills, N.Y.</td>
<td>Saul Geffner</td>
<td>A</td>
</tr>
<tr>
<td>Gardena High</td>
<td>Gardena, Cal.</td>
<td>David Marsh</td>
<td>A</td>
</tr>
<tr>
<td>Sacred Heart High</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Helen, OSF</td>
<td>A</td>
</tr>
<tr>
<td>Joliet Township High</td>
<td>Joliet, Ill.</td>
<td>Homer Kline</td>
<td>A</td>
</tr>
<tr>
<td>St. Augustine High</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Gregory, OSF</td>
<td>B</td>
</tr>
<tr>
<td>University High</td>
<td>W. Los Angeles, Cal.</td>
<td>Orin Peyton</td>
<td>A</td>
</tr>
<tr>
<td>St. Kieran High</td>
<td>Berlin, N.H.</td>
<td>Sr. Fidelia, OSF</td>
<td>A</td>
</tr>
<tr>
<td>Mt. Alvernia High</td>
<td>Pittsburgh, Pa.</td>
<td>Sr. Aelred, OSF</td>
<td>B</td>
</tr>
<tr>
<td>St. Mary High</td>
<td>Monroee, Pa.</td>
<td>Sr. Roseann, OSF</td>
<td>B</td>
</tr>
<tr>
<td>Duncon Fletcher High</td>
<td>Jacksonville Beach, Fla.</td>
<td>Richard von Genick</td>
<td>B</td>
</tr>
<tr>
<td>Paxon High</td>
<td>Jacksonville, Fla.</td>
<td>Vernon Hendricks</td>
<td>B</td>
</tr>
</tbody>
</table>
All teachers involved in this study had at least a bachelor's degree in chemistry and sixty percent had a master's degree. All had experience in teaching chemistry for at least five years. These teachers were selected by the writer because of their active participation in the New England Association of Chemistry Teachers, the National Science Teachers Association, or the American Chemical Society, or because they were recommended by reputable members of these organizations.

**Conditions, subjects and procedures.** In September, packets were mailed to each teacher participating in the study. These packets contained:

1. General time schedule
2. Introduction to the inductive unit
3. Inductive unit outline for equation-balancing
4. Inductive unit laboratory guide sheets (1 per pupil)
5. Detailed inductive laboratory notes for the teacher
6. Teacher's check list of steps to be followed and materials to be returned.
7. School data sheets (for inductive classes)
8. Intelligence test score sheets (for inductive classes)
9. Detailed inductive class notes for the teacher
10. Symbolic Unit-Test on equation-balancing (1 per pupil)
11. Answer sheets for the unit-test (2 per pupil)
12. Introduction to the deductive unit and unit outline
13. Deductive class notes for the teacher
14. Deductive laboratory notes (1 per pupil)
15. School data sheet (for deductive classes)
16. Intelligence test score sheet (for deductive classes)
17. Symbolic Unit-Test on equation-balancing (1 per pupil)
18. Answer sheets for unit-test (2 per pupil)
Each unit, inductive and deductive, was designed for eight forty-minute class periods of teaching. Two forty-minute class periods were needed for testing, one before and one after the unit was taught, making a total of ten classes over two complete five-day school weeks for the class instruction of this unit. Both inductive and deductive laboratory experiments were designed for five double or ten single laboratory periods.

The detailed class notes, given to each teacher to follow, strongly limited the variations possible in presenting the unit. The demonstrations permissible, the discussions to be held, and the information to be conveyed, were all carefully and clearly delineated, hence, the writer feels that this unit was handled in as closely similar a manner as is possible by more than forty individuals.

When the same teacher handled more than one chemistry class, she was asked to teach one by the inductive and one by the deductive method, using the notes supplied her by the writer for all the work. When the same teacher taught both an inductive and a deductive class, the personal variation was more restricted than in those cases where two teachers were involved. In the statistical treatment of the resultant data, the schools belonging to this first category are listed as Type A schools, whereas the weaker statistical setups in the discussion are called Type B schools. Throughout the study, the deductive classes are considered in the discussion as control groups, and the inductive classes as experimental groups.

The problem involved in Phase II of this study may be stated by a
null hypothesis: Teaching which includes inductive laboratory learning produces no statistically significant differences in the understanding of chemical equations as compared with deductive-descriptive learning and traditional teaching.

These differences were measured by the scores on a unique writer-constructed test which employs symbolic problems involving chemical reactions.

Method of approach.—Each teacher of an inductive class received a general introduction to inductive chemistry teaching in which the basic theory of the inductive approach to equation-balancing was given. Students taught by the inductive method see the reality in the laboratory before they learn the symbolic representation of it in class; they proceed from the tangible, from the observable and the known, i.e., known and observed in the experience of the student, to the problem the solution of which is unknown. The deductive method, on the other hand, proceeds by starting with a solution rather than a problem.

Secondly, each teacher participating in this study was asked to make notes of such comments. Pupil-teacher goals for the unit on equation-balancing were set up and printed for all teachers in the inductive unit outline. The goals were seven in number and are reproduced here. 1/

Pupil-teacher goals.—

1. The student should learn concepts through organized thinking, and the concept that is being applied here is the one of equation-balancing.

1/ See Appendix F.
2. The student should develop some **problem-solving ability** as in the preparation of salts by various methods.

3. The student should learn those particular **skills** which are specific to chemical laboratory work in testing the reactivity of substances.

4. The unit should make laboratory work particularly interesting and should increase observational powers and call attention to hidden meanings to be discovered in laboratory experimentation.

5. Some students may see the **practicality**, versatility, and great future of chemistry and be interested enough to make chemistry a vocational choice.

6. A greater **appreciation** of industrial chemistry and its impact on modern living should result.

7. A better **understanding** of the economy of the nation and its relation to chemistry will result from:
   a. Reading current periodicals.
   b. Study of industrial processes, particularly of gaseous reactions.
   d. Study of hydrocarbon combustions.

Core activities.-- These were required of all students and were clearly listed in the same unit outline. (See Appendix F.)

1. Read the assigned chapters on equation-balancing, reaction classification, solubility, and equilibrium in your textbook.

2. Listen to lectures given by the instructor on:
a. Laws of conservation of matter and of energy
b. Meaning of equations
c. Equilibrium: Law of Mass Action
d. Law of Le Chatelier

3. Do laboratory work without detailed and specific directions. Ideas and suggestions may be gleaned from outside reading.
   a. E.M.F. Series
   b. Synthesis and analysis
   c. Single and double replacement
   d. Reactions that go to completion
   e. Four ways to prepare a salt


The lectures and laboratory work referred to under these core activities were written out in great detail for each teacher and were included in the packet of material each received. (See Appendix F) If only these core activities were done, the unit would have been covered adequately and students could confidently attempt a test on equation-balancing concepts.

Related activities. — These were suggested for those students who were capable of more than a minimum amount of work and whose interest and initiative were high. The related activities ranged from book and periodical reading reports to field trips and laboratory experiments. The method of reporting the related activities was somewhat variable. The teacher could request either verbal or written accounts of the activity depending on her local situation. However, each activity that a student signed up to undertake had to be checked on by some means.

Evaluation. — A standardization of evaluative methods to determine
the degree to which students acquired desirable skills and behaviors is often somewhat difficult. The writer has established a method for standardizing evaluation of the degree of success or failure of the unit on equation-balancing. By listing nine headings and by giving sample replies for each heading, workable evaluative evidence was obtained. The page sent to the teachers contained the following headings and examples:

1. The skill and knowledge shown in the laboratory, aside from any mark on the actual accuracy of testing and identification, can evaluate general ability and intelligence.

SAMPLE: Sally G. keeps her work-area spotless; George C. has broken seven test tubes and three flasks so far.

A teacher can indicate by such comments just what type of laboratory worker a student proves to be. Skill in class problems must be accompanied by laboratory manipulative skill if the pupil is to make a scientist in the future.

2. The interest shown in laboratory work, the attentiveness, the observations, questions, inventiveness, all will indicate the students' approach to a scientific attitude.

SAMPLE: John W. was in the laboratory to try out an experiment with KMnO₄ which he read about while looking up material on sulfites.

Apparantly John W. would have some scientific curiosity which sets him apart from other students in the same class who read the same references and did not have the same interest which he had. A further notation on how John carried out his testing with the permanganate would tell a little about his inventiveness, and perhaps his powers of observation.
3. The problem-solving part of a unit in which students must devise, carry out, and explain laboratory procedures will tell much about their ability to reason.

SAMPLE: Malcolm F. has a splendid notebook with detailed observations, conclusions, and references.

If a student can use a reference well, can assimilate the directions and information contained in several references, and can write up an experiment intelligently as an integrated whole, that student has skills which are fundamental to problem-solving and which would be strong tools for the development of a future scientist.

4. The chemicals and equipment the students request are a good indication of their growth in and assimilation of chemical knowledge.

SAMPLE: Roger P. requested six additional metals to further test the E.M.F. series.

Students may refuse to use the chemicals suggested for a particular experiment and may request material to suit their fancy. Often this is a bid for mere time-wasting, as the good student generally first uses what is supplied then, if more testing seems to him to be in order, he may ask for more equipment and chemicals. The type of material he seeks and the use he makes of it are good measures of his degree of chemical knowledge at the time. An extraordinary student may occasionally be found who already knows the elementary material of the unit. Teachers will recognize such a pupil immediately and would allow the use of chemicals and equipment for more advanced study. Originality should never be penalized, but encouraged.

5. Use of references in correlation with observations indicates
ability to synthesize the material gathered in reference to the point under investigation.

SAMPLE: George P. turned to the Handbook of Chemistry and Physics to determine the relative solubilities of two cadmium salts.

When a student can turn quickly to a source book and discover the information he needs, he has acquired a proficiency in and acquaintance with chemical literature that is a mark of maturity in a high school science student.

6. The type of reference book used and the student's interpretation of it is a clear indication of his level of learning.

SAMPLE: Maureen V. used Kolthoff and Sandell's book on quantitative analysis for her related activity experiment; John B. used a high school laboratory manual for his.

Use of a deductive high school manual of chemistry for extra work means little but that a student can follow directions, whereas, the intelligent use of a college text means that the student is capable of pushing further than the limited high school course, and has some initiative and scientific interest.

7. The final report sheets will show the degree to which the material has been assimilated.

SAMPLE: James W. passed in untidy and unorganized report sheets. A student who is not orderly in his written work reports, who lacks neatness and precision, can be spotted in high school with little trouble. The time to direct or correct him is then, before his final choice is made and he has embarked on the wrong college course. The method is far from infallible, as the late maturing pupil is ever
present, yet, observation of this power to assimilate can be a useful guidance implement.

8. The procedure used in the laboratory will indicate and differentiate between the leaders and the followers, the thinkers and the technicians.

SAMPLE: Betty T. waits for Mary S. to do an experiment and then she imitates it.

The followers can easily be detected, as also can those with initiative, ideas and courage. Those who really know what they are doing stand out from the group who merely copy. When original experiments are given, this is much more apparent than when all are blindly following a set of directions.

9. The approach to the laboratory problem of preparing a salt in four ways will indicate the amount of understanding the student has of what has gone before.

SAMPLE: Vivian K. merely copied directions from a borrowed laboratory manual; George P. made four salts instead of one.

This experiment provides the clearest indication of what the student has learned from this unit. If he makes one salt by four methods, purifies his products and presents good, dry crystals, he has mastered the topic. If he makes four salts instead of one, but does them well, he has acquired good technical skill; if he merely takes old directions from deductive manuals and applies them, he is lacking in that ability to think chemically which the inductive laboratory learning was supposed to instill.

With this list of headings in mind, a teacher can watch more
intently and evaluate more critically from day to day so that her final evaluation of the unit as a whole would not so easily deteriorate into a snap judgment or perfunctory remark.

2. The Inductive Class

Notes supplied to each participating teacher were written for eight forty-minute class periods and covered four major headings related to chemical equation-balancing. (See Appendix F.)

A typical class reaction.-- The first day's lecture (Appendix F) was an introduction to the unit, a setting up of the problem to be solved. Most of the discussion had to be elicited from the students themselves. They had to be made to feel a vital interest in the work and to come to the realization that it held the key to problems they had previously met but could not solve. The majority had never come to grips with these problems because they had never delved beneath the surface in their thinking. But the class discussion immediately started them probing. The simple example of burning wood was given. Would the products weigh the same as, more than, or less than the wood? What reacted with the wood? Has it weight? Have all the products weight? Even the slowest student was brought into such a discussion and contributed to it. Once the concept of combination and weight equality was firmly established the laws of conservation were enunciated. They were then laws with meaning— not empirical statements to be taken and memorized because a book had printed them in italics. The student had proceeded from a known fact to an unknown— he had solved a problem inductively. Numbers were then introduced (preferably in the o.g.s. system) and a digression into atomic reactions was made if the class
discussion tended that way. It was presumed that equivalence of matter and energy according to the Einsteinian viewpoint would be dealt with thoroughly in a unit on nucleonics, and hence it was not specifically included in the topic in hand.

The second and third lectures treated symbolic equations and were not undertaken until direct combination and analysis reactions had been performed in the laboratory. The student had to have this experience to build upon; otherwise the inductive character of the work would have been lost. The first of these two lectures was spent principally in a demonstration of electrolysis of water as a basis from which to write the equation for water decomposition. The concept of conservation of matter and the additive property of weight and the non-additive property of volume were clarified. A schematic diagram illustrative of steam decomposition was included and served to concretize this abstract idea.

The third topic, equilibrium, embraced three classes, the first being devoted to ionic equilibria and the gaseous equilibrium of the Haber process. The second class treated of the variables which can alter the equilibrium point, particularly the concentration. Not until the third class was it generally possible to elicit from the students a principle which explained the equilibrium phenomena. Then and only then could the law of Mass Action be stated.

The fourth and last topic in the unit was Le Chatelier's Principle on which two more class periods were spent. Temperature and pressure
changes were considered in relation to the Haber Process. A great deal of student participation was required here to get this matter clear and straight. Another schematic diagram was suggested for simplifying the understanding of pressure effects on the closed system of equilibrium. (See Appendix F.)

The last day was spent in a general discussion of the unit and in practice drill on equation balancing. Seventy-five equations (Appendix F) were included in this unit and these were done by the students in class or at home study throughout the unit.

3. Inductive Laboratory

The packets of materials sent to the teachers included a set of laboratory guide sheets for each student. These sheets gave suggestions of what was to be done in the laboratory, but offered no help on how to accomplish this work. For the early October experiment on "Types of Reactions," students were told that four types existed: analysis, synthesis, single and double replacement, but instead of the traditional directions found in ordinary high school manuals, the student read:

"Collect the necessary materials. You are asked to perform two experiments of each type listed above. Identify all products of reaction. You will have to consult your textbook and auxiliary texts to find these tests. Report the methods you applied and give complete balanced equations for every test and reaction used, including those in testing for ions, elements, or radicals."

To accompany this guide sheet, the teacher received detailed laboratory notes to prepare her for what might happen once the students started testing. The synthesis reactions and the tests involved were allowed to occupy seventy-five minutes of laboratory time and were

See Appendix F.
generally taken first since this section is the foundation for the understanding of the meaning of reactions. The teacher was instructed to supply students with magnesium ribbon, red phosphorus, powdered and lump charcoal, iron filings, powdered and roll sulfur. The materials are safe, easy to handle, and simple to use, hence, for early in the course they are ideal.

A typical class reaction.— Once a student had burned a piece of magnesium it became very evident to him that the product, white magnesium oxide powder, differed from the reactants, silvery magnesium metal and colorless oxygen gas. The teacher played an important part throughout the inductive laboratory sessions, and particularly so in the beginning phases of experimentation. By questioning, by drawing analogies, by working gradually from the known and observed experience of the students she had to cause an awareness and comprehension of the true nature of the change involved in the burning of magnesium and of other elements. The equation finally set down by the pupil was probably not balanced unless the Law of Conservation of Matter had already been taught in class. It was considered advisable to leave the equation unbalanced at first, so when a student put on paper:

\[ \text{Mg} + \text{O}_2 \rightarrow \text{MgO} \]

the matter was considered closed for the time and he was permitted to start working with another element such as sulfur. The sulfur equation came more quickly to him. He was not expected to identify MgO, but SO$_2$ he might have used to decolorize permanganate. Some had not that skill and hence were not forced to do this additional testing. The odor of
sulfur dioxide was remarkable by all. However, when CO₂ was prepared by the burning of carbon, each student was required to perform a positive identification test. The students used these elements in any order and no restriction was placed on their choices. Before the laboratory period was over, teachers made all students check the reactions they had written to see that the number of atoms on both sides of the arrow was equal.

For the analysis experiment, another seventy-five minutes was allotted. Teachers were asked to supply mercuric oxide and potassium chlorate for student use. Most students proceeded to heat these substances, some in test tubes, and others on spatula-tips. The better students tested O₂ immediately; the slower ones responded after questions or suggestions had been made by the teacher, or after they had seen what a neighboring student had done. The leaders and the followers began to be differentiated as early as this in the unit.

The ternary salt, K₂O₃, presented difficulties to many pupils, some of whom still had not grasped the diatomic character of oxygen; others thought analysis invariably broke compounds into elements. By insisting on complete identification of products, teachers brought students face to face with the positive evidence for the existence of potassium and chloride ions, and oxygen gas. In this way valuable personal experience was had. In classes where time permitted, some more synthesis reactions were performed. These differed from the previous work in that oxygen was not one of the reactants. A better and broader concept of synthesis resulted from this experimentation.
Single replacement reactions and electromotive series experimentation were considered completed by one seventy-five minute session on the E.M.F. Table. Students were given a list of metals and acids to use, including sodium, calcium, magnesium, aluminum, copper, zinc, iron, mercury, tin, cadmium, and lead; dilute and concentrated acetic, nitric, hydrochloric and sulfuric acids. Most students were quite content with this array of chemicals; some wanted extras once they had used these; others wanted replacements for these. Generally speaking, the former requests reflected a grasp of the theory involved, while the latter indicated either that the student was slavishly following "recipes" he had copied from books, or, that he had no organized plan in mind and merely wanted to waste time. Teachers were asked to check constantly to see that students really were identifying their chemical products, particularly the gaseous ones. Slower students worked heterogeneously on metals and acids; brighter and more disciplined minds set about the matter systematically and used one metal with several acids and several metals with one acid.

Teachers were asked to check the forms of the metals provided to be sure they were all shot, or all powder, or all strip. Students had not learned the factors influencing reaction rates and so would have been needlessly confused if variation in particle size had been introduced. An occasional student tested the replacement of a lower metal by one higher in the series, and encouragement was given to any who probed that far.

The double displacement reactions met with greater success if performed after the set on "Reactions that go to completion." Seventy-
five minutes was again allotted but the major portion of the time was, in most cases, occupied by the teacher in directing discussions on the topic. Students were encouraged to pair up reagents from the general reagent shelf and see what evidence they could obtain for reaction.

By the time double displacement reactions were reached, the students had acquired much more skill and confidence. They paired up solutions they had used before, such as a chloride and silver nitrate, and obtained silver chloride as a precipitate. Sulfates plus barium chloride gave them a barium sulfate precipitate. By degrees the students learned soluble from insoluble compounds. Some failed to make this transition and were spotted by the teachers with the greatest of ease.

The last experiment treated of soluble and insoluble salts and asked students to prepare two of each and in addition to make sodium chloride and barium sulfate by four distinct processes. The manner in which the students attacked this experiment clearly indicated their grasp of chemical reactivity and reaction balancing. The slow student did very little; the memorizer repeated what he had done previously but could not abstract from prior concrete situations the theory which would carry him to success in this venture. The regular student did fairly well, but had impure salts as products; the superior pupil produced neat and finished work, dried his products and showed complete mastery of the topic.

4. Test

The Symbolic Unit-Test administered before and after the teaching
of the unit was designed by the writer to measure understanding of the concept of equation-balancing. Instead of completely verbal questions, the answers to which might be rote memorization, the questions were combined with diagrammatic sketches of laboratory setups. The relationship of laboratory work to equation-balancing was thus stressed. Too often the two are divorced in presentation and never come together in the mind of the child. The answers to the questions were not difficult, nor were the questions intended to be subtle. They involved a higher level of thinking than would a true-false or completion test, and were principally of the type:

a. An increase in A will produce an increase in B.
b. An increase in A will produce a decrease in B.
c. An increase in A will have no effect upon B.

and

da. The reason given is true and applies to the question.
b. The reason given is true and does not apply to the question.
c. The reason given is false.
CHAPTER VII

ANALYSIS OF DATA - PHASE II

1. Choice of Statistical Tools

In Phase II of the study statistical tools had to be chosen for the analysis of the data collected by the administration of the Symbolic Unit-Test on equation-balancing. (Appendix F)

In this case, the tool selected had to have sensitivity sufficient to test the difference between means if true differences really existed; it had to be able to test differences in groups of varying size; and lastly, it had to be capable of adjusting data to the criteria means when the control variations affected the criteria.

2. Treatment of Data

To test the null hypothesis that no true differences exist between inductively taught classes and deductively taught classes in terms of the Symbolic Unit-Test administered on the matter of the equation-balancing unit, a critical ratio could be determined. However, if the criteria means are affected by the mental ability or the pre-test scores, inferences from critical ratios could prove erroneous. Hence, the intelligence test scores 1/ of all Type A

1/ Otis Quick Scoring Mental Ability Test, Gamma Form, World Book Company, Yonkers-on-Hudson, New York, 1937.
schools were used to determine critical ratios of intelligence between inductive and deductive classes. The same was done for the Type B schools. Finally a critical ratio of mental ability was established for all inductively taught classes and for all deductively taught classes. This was done in order to uphold or reject the null hypothesis that all classes being used in the study are of comparable mental ability as measured by an intelligence test. The results of this work appear in Table 27.

Table 27. Critical Ratios of Mental Ability Test Scores, Type A Schools

<table>
<thead>
<tr>
<th>Type Class</th>
<th>Mean IQ</th>
<th>S.E.</th>
<th>Difference</th>
<th>S.E.D.</th>
<th>O.R.</th>
</tr>
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<tbody>
<tr>
<td>Deductive</td>
<td>103.0</td>
<td>1.75</td>
<td>0.30</td>
<td>2.5</td>
<td>.12</td>
</tr>
<tr>
<td>Inductive</td>
<td>107.7</td>
<td>1.87</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

When a critical ratio is 1.96 or more, a null hypothesis may be rejected at the five percent level of confidence. In the above case the critical ratio falls far short of this figure, therefore, the null hypothesis is retained and there is no statistical difference in mental ability between inductively and deductively taught classes in Type A schools.

Table 26. Critical Ratios of Mental Ability Test Scores, Type B Schools

<table>
<thead>
<tr>
<th>Type Class</th>
<th>Mean IQ</th>
<th>S.E.</th>
<th>Difference</th>
<th>S.E.D.</th>
<th>O.R.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>108.41</td>
<td>1.71</td>
<td>.72</td>
<td>2.22</td>
<td>.32</td>
</tr>
<tr>
<td>Inductive</td>
<td>107.69</td>
<td>1.42</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[\text{Otis Quick Scoring Mental Ability Test, op. cit.}\]
Again, in the case of the Type B schools the critical ratio falls far short of the value 1.96 which would cause the null hypothesis to be rejected at the five percent level of confidence, hence, the null hypothesis is again retained and it is assumed that there is no appreciable difference in mental ability between the inductive and the deductive classes in the Type B schools.

Table 29. Critical Ratios of Mental Ability Test Scores, Type A and Type B Schools, Inductive and Deductive Classes Pooled

<table>
<thead>
<tr>
<th>Type of Class</th>
<th>Mean IQ</th>
<th>S.E.</th>
<th>Difference</th>
<th>S.E.</th>
<th>C.P.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>103.11</td>
<td>1.59</td>
<td>-4.1</td>
<td>2.11</td>
<td>.169</td>
</tr>
<tr>
<td>Inductive</td>
<td>107.7</td>
<td>1.39</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The critical ratio is again too small for the rejection of the null hypothesis so it is assumed that the classes of the study are all similar in mental ability as measured by the Otis Intelligence test. 1/ The similarity in intelligence test scores from such widely scattered schools may well be accounted for by the fact that chemistry classes under consideration are all of the college preparatory variety and hence might be expected to have the better students enrolled in them.

The means and standard deviations were found for the pre-test scores of each class in Type A and Type B schools. The standard deviations of these scores were compared by the use of the F-test, a ratio of the larger variance to the smaller. This F-test furnishes

1/ Otis Quick Scoring Test of Mental Ability, op. cit.
an overall comprehensive test of the significance of differences
between means. A significant F does not tell which means differ
significantly, but at least that one is reliably different from
some others. Since the F was not significant in any case, none of
the mean differences were significant and on these grounds, together
with the critical ratios already described, it was assumed that
inductive classes of the Type A schools and inductive Classes of the
Type B schools did not differ significantly in either general ability
or in previously acquired chemical knowledge and hence could be
grouped together or pooled. The deductive classes of Type A and Type
B schools were treated in a similar manner and also showed no
significant variation and so were grouped for study.

Having thus determined that aggregate groups of inductive and
deductive classes were of comparable intelligence, a form of t-test
was chosen to test the null hypothesis of the equivalence of the
inductive and deductive methods in teaching the unit on chemical
equation-balancing.

Because the post-test is the repetition of the same measurement
as the pre-test, what is desired is a comparison of the mean gains in
the two groups, inductive and deductive. When means of such paired
samples are not independent, but correlated as in this case, the best
formula to use for deriving the t-value is the following: 1/

\[
  t = \frac{M_I - M_D}{\sqrt{\frac{S^2}{N(N-1)}}}
\]

In which \( M_I \) = mean gain of inductive group from pre- to post-test

\( M_D \) = mean gain of deductive group from pre- to post-test

Where

\[ \varepsilon d^2 = \varepsilon D^2 - \frac{(\varepsilon D)^2}{N} \]

Where

\[ \varepsilon D^2 = \varepsilon X^2 + \varepsilon Y^2 - 2 \varepsilon XY \]

And

\[ \varepsilon D = \varepsilon X - \varepsilon Y \]

\( X \) = pre-test scores

\( Y \) = post-test scores

\( N \) = total number of students involved

The t-formula above is based on the general \( t = D/SE_D \), but in this case the denominator is a less familiar form for determining the standard error.

In Type A schools the data was processed first. Using the scores of the Symbolic Unit-Test (Appendix F) administered before and after the teaching of the unit on equation-balancing by the inductive and deductive methods, the t-value was calculated in the following manner:

There were 425 inductive and 425 deductive pupils, hence \( N = 850 \).

The summation of the pre-test scores for deductive classes, \( \varepsilon X = 4764 \).

The summation of the pre-test scores for inductive classes, \( \varepsilon X_1 = 4939 \)

The summation of the post-test scores for deductive classes, \( \varepsilon Y = 6628 \)

The summation of post-test scores for inductive classes, \( \varepsilon Y_1 = 8553 \)

Hence, total pre-test scores: \( \varepsilon X = 9703 \)

Total post-test scores: \( \varepsilon Y = 15181 \)

By formula, the difference in these summations is \( \varepsilon D \):

\[ \varepsilon D = \varepsilon Y - \varepsilon X; \quad \varepsilon D = 15181 - 9703 = 5478 \]
The sum of squares for all deductive class pre-test scores: \( \leq X^2 = 61751 \)

The sum of squares for all inductive class pre-test scores: \( \leq X_i^2 = 66108 \)

The total sum of pre-test scores: \( \leq X = 127859 \)

The sum of squares of all deductive post-test scores: \( \leq Y^2 = 102456 \)

The sum of squares for all inductive post-test scores: \( \leq Y_i^2 = 174046 \)

The summation of the products of pre-test and post-test scores:

\[ \leq XY = 73827 \]

\[ \leq X_i Y_i = 100749 \]

From this data the summation of the squared differences is found:

\[ \leq D^2 = \leq X^2 + \leq Y^2 - 2 \leq XY \]

\[ \leq D^2 = 404361 - 349152 \]

\[ \leq D^2 = 55209 \]

This number minus the correction factor \( \frac{(\leq D)^2}{N} \) gives \( \leq d^2 \),

\[ \leq d^2 = 55209 - \frac{(5478)^2}{850} \]

\[ \leq d^2 = 19905 \]

With this much information the t-formula can be applied since the mean scores of both groups are known.

\[ t = \frac{M_I - M_D}{\sqrt{\frac{\leq d^2}{N(N-1)}}} \]

\[ t = \frac{9.0 - 4.7}{\sqrt{\frac{19905}{850(49)}}} \]

\[ t = \frac{4.3}{17} \]

\[ t = 25.5 \]
Consulting a table of t-values using degrees of freedom equal to the number of scores tallied, or 850, t is significant at the one percent level of confidence if its value is 2.58 or greater. Below in table form are the results of the thirty-six classes scored at the time the paper was written.

Table 30. Application of the t-test to Pre-Test — Post-Test Score Gains on the Symbolic Unit-Test, Type A Schools, Inductive and Deductive Classes

<table>
<thead>
<tr>
<th>Score Summations</th>
<th>Deductive (1)</th>
<th>Inductive (2)</th>
<th>Formula Applications (3)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \Sigma X )</td>
<td>4751</td>
<td>4993</td>
<td>( \Sigma D ) 5333</td>
</tr>
<tr>
<td>( \Sigma Y )</td>
<td>6470</td>
<td>8607</td>
<td>( \Sigma D^2 ) 57371</td>
</tr>
<tr>
<td>( \Sigma X^2 )</td>
<td>61269</td>
<td>66472</td>
<td>( \Sigma d^2 ) 23912</td>
</tr>
<tr>
<td>( \Sigma Y^2 )</td>
<td>103218</td>
<td>174516</td>
<td>( \Sigma d ) 4.0</td>
</tr>
<tr>
<td>( \Sigma XY )</td>
<td>73175</td>
<td>100747</td>
<td>( \Sigma I ) 8.4</td>
</tr>
<tr>
<td>( N )</td>
<td>425</td>
<td>425</td>
<td>( t ) 31.4</td>
</tr>
<tr>
<td>( N_{\text{pre}} )</td>
<td>11.2</td>
<td>11.7</td>
<td></td>
</tr>
<tr>
<td>( N_{\text{post}} )</td>
<td>15.2</td>
<td>20.3</td>
<td></td>
</tr>
</tbody>
</table>

From the foregoing data the t-value is evidently significant at the one percent level of confidence and it is justifiable to reject the null hypothesis which declares that inductive and deductive methods are equally efficient in teaching the concepts of equation-balancing as measured by the Symbolic Unit-Test. The data support the belief that the inductive method is superior at the one percent level of confidence in Type A Schools.

Next the Type B schools were studied. The results appear in Table 31.

\( \text{\textsuperscript{1}} \) Henry Garrett, op. cit. p. 429.
Table 51. Application of the t-Test to Pre-Test--Post-Test Score
Gains on the Symbolic Unit-Test, Type B Schools,
Inductive and Deductive Classes

<table>
<thead>
<tr>
<th>Score Summations</th>
<th>Formula Applications</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>Inductive</td>
</tr>
<tr>
<td>(1)</td>
<td>(2)</td>
</tr>
<tr>
<td>Σx ... 2683</td>
<td>Σx₁ ... 1991</td>
</tr>
<tr>
<td>Σy ... 2644</td>
<td>Σy₁ ... 5555</td>
</tr>
<tr>
<td>Σx² ... 27591</td>
<td>Σx²₁ ... 24369</td>
</tr>
<tr>
<td>Σy² ... 32512</td>
<td>Σy²₁ ... 70696</td>
</tr>
<tr>
<td>Σxy ... 51304</td>
<td>Σxy₁ ... 39769</td>
</tr>
<tr>
<td>n .... 100</td>
<td>n₁ ...... 100</td>
</tr>
<tr>
<td>H₁pre ... 11.4</td>
<td>H₁pre ..... 11.1</td>
</tr>
<tr>
<td>H₁post ... 14.7</td>
<td>H₁post ... 19.9</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

With 360 degrees of freedom a t-value of 2.59 is significant at
the one percent level of confidence and it is justifiable to reject
the null hypothesis which asserts that inductive and deductive methods
are equally efficient in teaching the concepts of equation-balancing.
The data support the belief that the inductive method is superior to
the deductive-descriptive method at the one percent level of confidence
as measured by the Symbolic Unit-Test on equation-balancing.

The t-test was then applied to the scores of all inductive classes
and all deductive classes on the Symbolic Unit-Test. The inductive
classes of both Type A and Type B schools were pooled as were the
deductive classes of both types of schools since these classes have
been shown to be similar in mental ability and in prior chemical
knowledge. The results of the t-test appear in Table 52.
Table 32. Application of the t-test to Pre-Test—Post-Test Score Gains on the Symbolic Unit-Test, Inductive and Deductive Classes of Type A and Type B Schools Pooled.

<table>
<thead>
<tr>
<th>Score Calculations</th>
<th>Formula Applications</th>
</tr>
</thead>
<tbody>
<tr>
<td>Inductive (1)</td>
<td>Deductive (2)</td>
</tr>
<tr>
<td>$X_{1**}$ 6989</td>
<td>$X$ ** 6016</td>
</tr>
<tr>
<td>$X_{2**}$ 12192</td>
<td>$X^{2}$ ** 9114</td>
</tr>
<tr>
<td>$X_{1**}$ 90841</td>
<td>$X^{2}$ ** 88790</td>
</tr>
<tr>
<td>$X_{1**}$ 245002</td>
<td>$X^{2}$ ** 142530</td>
</tr>
<tr>
<td>$X_{1**}, Y_{1**}$ 140527</td>
<td>$X_{1**}Y_{1**}$ 164559</td>
</tr>
<tr>
<td>$N$ 605</td>
<td>$N$ ** 605</td>
</tr>
<tr>
<td>$X_{pre}$ 11.5</td>
<td>$X_{pre}$ 11.2</td>
</tr>
<tr>
<td>$X_{post}$ 26.1</td>
<td>$X_{post}$ 15.0</td>
</tr>
</tbody>
</table>

Therefore, again at the one percent level of confidence a significant difference exists between inductive and deductive classes and the null hypothesis regarding their equality is rejected. Apparently there is sufficient evidence, on the basis of scores on the Symbolic Unit-Test on equation-balancing that the inductive method is superior in importing the chemical concepts involved in equation-balancing.

At the time the foregoing tables were constructed some classes had not reported their data. Those received later may be found in Appendix G.

The following work on quartiles was also completed with some classes missing. When the material arrived from these classes it was tallied and added into the data for Table 32 and succeeding tables.

*Weighted mean. See Chapter IV.
final results were not significantly different from those shown thereby signifying that the sample used gave reliable results.

The effects of high and low intelligence upon results obtained by inductive and deductive learning were then investigated. The upper and lower quartiles of the first thousand papers scored (500 inductive and 500 deductive) were selected thus making 125 in each group, inductive and deductive, for each quartile determination. The means-standard deviations of these papers were determined and the results for the upper quartile intelligence groups appear in Table 33.

Table 33. The Means-Standard Deviations, Upper Quartile Intelligence, Inductive and Deductive Classes, Symbolic Unit Test Scores

<table>
<thead>
<tr>
<th>Class Type</th>
<th>Post-Test</th>
<th>Pre-Test</th>
<th>Difference</th>
<th>S.D.</th>
<th>N</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>13.4</td>
<td>11.6</td>
<td>1.8</td>
<td>2.6</td>
<td>125</td>
</tr>
<tr>
<td>Inductive</td>
<td>16.0</td>
<td>11.6</td>
<td>4.3</td>
<td>1.7</td>
<td>125</td>
</tr>
</tbody>
</table>

The question then arises: is the gain from pre-test to post-test score significant in both the inductive and the deductive groups? The experimental design called "single group method" was first used in this calculation. In order to determine the significance of the difference between means obtained in the initial and final testing, the standard error of the difference must first be calculated. This was done by means of the formula for the standard error of the difference between correlated means:

$$\text{S.E}_{D^1} = \sqrt{\frac{\sigma^2_{M_1}}{n_1} + \frac{\sigma^2_{M_2}}{n_2} + 2\rho_{12} \sigma_{M_1} \sigma_{M_2}}$$

1/ Otis Quick: Scoring Test of Mental Ability, op. cit.
In which $M_1$ and $M_2$ are the standard errors of the initial and final means and $r_{12}$ is the coefficient of correlation between scores made on initial and final tests.

$$
\begin{array}{|c|c|c|}
\hline
\text{Pre-Test} & \text{Post-Test} \\
\hline
\text{Number} & 125 & 125 \\
\text{Mean score} & 11.6 & 13.4 \\
\text{Standard dev.} & 2.4 & 2.6 \\
\text{S.E.}_M & .48 & .52 \\
\hline
\end{array}
$$

Difference between means \[ 1.8 \]

Coefficient of correlation \[ .51 \]

$$
S_{E_D} = \sqrt{(.48)^2 + (.52)^2 - 2(.51)(.48)(.52)}
$$

$$
S_{E_D} = .49
$$

$$
t = \frac{D}{S_{E_D}} = 1.8 / .49 = 3.67
$$

As there are 125 students, there are 125 pairs of scores, so the degrees of freedom become 125-1 as one degree of freedom is lost since the $S_{E_D}$ is computed around the distribution of the differences. From the tables of $t$-values $^{1}$ with 124 degrees of freedom, $t_{.01} = 2.62$. The $t$-value calculated is greater than that and is therefore significant at the one percent level of confidence. It is evident that the classes made substantial progress from initial to final test scores. In treating the inductive classes, a similar test was made and the following figures were determined:

$$
\begin{array}{|c|c|c|}
\hline
\text{Pre-Test} & \text{Post-Test} \\
\hline
\text{Number} & 125 & 125 \\
\text{Mean score} & 11.6 & 16.0 \\
\text{S.D.} & 2.2 & 1.2 \\
\text{S.E.}_M & .44 & .24 \\
\hline
\end{array}
$$

$$
\text{Difference between means} \[ 4.4 \]
$$

Coefficient of correlation \[ .51 \]

$^{1}$ Henry Garrett, op. cit., p. 429.
Again the t-value is significant at the one percent level of confidence and the gain made by the inductive group was significant.

The difference method was also used on this and the succeeding calculations. The results appear in Table 34. In this method t the standard error of the difference is calculated from the formula:

\[ S.E_{MD} = \frac{S.D_D}{\sqrt{N}} \]

Table 34. Significance of Post-Test Unit-Test Scores, Symbolic Unit-Test, Inductive and Deductive Classes, Upper Quartile IQ

<table>
<thead>
<tr>
<th>Class Type</th>
<th>Score Difference</th>
<th>S.D.</th>
<th>S.E.</th>
<th>S.E. MD</th>
<th>t</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>1.8</td>
<td>2.8</td>
<td>.25</td>
<td>7.2</td>
<td></td>
</tr>
<tr>
<td>Inductive</td>
<td>4.3</td>
<td>1.7</td>
<td>.15</td>
<td>23.7</td>
<td></td>
</tr>
</tbody>
</table>

With 125 degrees of freedom the t-value of 2.62 is significant at the one percent level of confidence, hence, it is evident that there is a significant difference between pre-test and post-test scores at the one percent level of confidence in both inductive and deductive classes.

Since the pre-test scores in both classes were alike, the difference method was applied to post-test scores in order to determine the t-value and thus retain or reject the null hypothesis: there is no significant difference in achievement between inductive and deductive

\[ \text{I/ Henry Garrett, op. cit. p} 227. \]
methods of teaching the upper quartile intelligence students as measured by the Symbolic Unit-Test on chemical equation-balancing.

The results of these calculations appear in Table 35.

Table 35. Application of the t-Test to the Symbolic Unit-Test Post Test Scores, Upper Quartile Intelligence, Deductive and Inductive Classes

<table>
<thead>
<tr>
<th>Type Class</th>
<th>Mean</th>
<th>S.D.</th>
<th>N</th>
<th>S.D. M</th>
<th>S.E. D</th>
<th>t</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>15.4</td>
<td>2.8</td>
<td>125</td>
<td>2.2</td>
<td>1.08</td>
<td>13.6</td>
</tr>
<tr>
<td>Inductive</td>
<td>16.0</td>
<td>1.7</td>
<td>125</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

A t-value of 2.58 is significant at the one percent level of confidence, hence, this value for t is significant and the null hypothesis is rejected. The students of higher intelligence benefit more by inductive teaching than do the students deductively taught, as measured by their achievement on the Symbolic Unit-Test on equation balancing.

The lower quartile intelligeace students were then considered.

Table 36. Means-Standard Deviations of the Lower Quartile Intelligence, Symbolic Unit-Test Scores, Inductive and Deductive Classes

<table>
<thead>
<tr>
<th>Type Class</th>
<th>Post-Test</th>
<th>Pre-Test</th>
<th>Difference</th>
<th>S.D.</th>
<th>N</th>
</tr>
</thead>
<tbody>
<tr>
<td>Deductive</td>
<td>11.5</td>
<td>10.6</td>
<td>0.7</td>
<td>2.7</td>
<td>125</td>
</tr>
<tr>
<td>Inductive</td>
<td>14.7</td>
<td>10.7</td>
<td>4.0</td>
<td>2.5</td>
<td>125</td>
</tr>
</tbody>
</table>

Again the question occurs: is the gain in both cases, inductive and deductive, significant? Application of the t-test by the difference
method, to inductive and deductive scores follows.

Table 37. Application of the t-Test to Symbolic Unit-Test Post-Test Scores, Lower Quartile Intelligence, Inductive and Deductive Classes

<table>
<thead>
<tr>
<th>Type Class</th>
<th>Mean</th>
<th>S.D.</th>
<th>N</th>
<th>S.D.</th>
<th>S.E.</th>
<th>t</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1)</td>
<td>(2)</td>
<td>(3)</td>
<td>(4)</td>
<td>(5)</td>
<td>(6)</td>
<td>(7)</td>
</tr>
<tr>
<td>Reductive</td>
<td>11.5</td>
<td>2.4</td>
<td>125</td>
<td>2.4</td>
<td>.22</td>
<td>14.5</td>
</tr>
<tr>
<td>Inductive</td>
<td>14.7</td>
<td>2.5</td>
<td>125</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Using 250 degrees of freedom as the closest figure with which to enter the tables \( 1/ \), the t-value of 2.59 is found to be significant at the one percent level of confidence. The t-value obtained by the calculations is therefore significant and the null hypothesis is rejected. Inductive teaching is significantly better than deductive teaching at the one percent level of confidence as measured by the Symbolic Unit-Test on equation-balancing. The lower quartile students inductively taught are superior to the lower quartile students deductively taught in their knowledge of chemical concepts of equation-balancing.

The results of the t-tests applied to the Symbolic Unit-Test scores indicate that the inductive method is statistically superior to the deductive-descriptive method with average classes, with students in the lowest quartile in intelligence, and with students in the highest quartile in intelligence, in the mastery of the concepts of chemical equation-balancing.

3. The Symbolic Unit-Test

Item Analysis. -- The proportion of success in the highest and lowest twenty-seven percent of the test scores obtained was used in

\(^1/\) Henry Garrett, op. cit., p. 429.
the determinations. Later a random hundred papers were chosen of the rest of the test papers and a similar analysis was made on that sample. The figures varied but slightly, and not significantly, from those given for the thousand papers, hence, these figures were retained as reliable.

The difficulty index for each of the thirty-six test items was desired. This difficulty index may be defined as the proportion of a certain sample of the testees that knows the answer to the item correctly, after correction has been made for chance. The discrimination index, which has also been determined for each item, is a measure of the discriminating power of the item. It may be stated as a critical ratio, a biserial moment of correlation, a product-moment coefficient of correlation, or the phi-coefficient. The critical ratio would indicate how certain we might be that a given item does discriminate between high and low scoring groups, but it does not give the amount of discriminating ability possessed by each item. The biserial and product-moment coefficients and the phi-coefficient are all affected by the percent of the sample correctly answering the test questions. This makes it difficult to assess the value of questions answered correctly by widely different proportions of the sample.  

The best measure of the relationship for internal consistency item analysis is the biserial coefficient of correlation between each item and the total score, exclusive of this item being studied. Items with discrimination indices above .20 are considered adequate for the

\[ \text{Frederick H. Davis, Item Analysis Data. Harvard Education Press, Cambridge, 1949.} \]
discrimination of students on this unit achievement test.

For each item of the Symbolic Unit-Test, the proportion of successes was computed by a method which allowed for corrections for guessing and chance. The general formula used was: \( P_H = \frac{R_H}{N_H - N_{RH}} \)

\[ P_H = \frac{R_H}{N_H - N_{RH}} \]

In which:

- \( N_H \) = number of testees in the upper 27 percent of the sample
- \( R_H \) = number of testees in the upper 27 percent of the sample who answered the test item correctly
- \( \overline{W}_H \) = number of testees in the upper 27 percent of the sample who answered the test item incorrectly
- \( N_{RH} \) = number of testees in the upper 27 percent of the sample who did not reach the item
- \( X \) = number of possible answers to the test item

This calculation was made on the upper twenty-seven percent of the first thousand papers scores, and also on the lowest twenty-seven percent of the papers. A random sample of 100 papers was later drawn from the remaining papers and the determinations made on them confirmed the above results.

The figures obtained from the application of the formula were then used to enter a table of discrimination and difficulty indices devised by Frederick Davis of Harvard University School of Education. This table accompanies his booklet on Item Analysis Data and was used throughout this section.

\[ P_{FH} \]

\[ \text{Frederick Davis, op. cit., enclosed chart} \]
Table 3B. Item Analysis Data for the Symbolic Unit-Test on Equation-Balancing

<table>
<thead>
<tr>
<th>Item No.</th>
<th>Difficulty Index (1)</th>
<th>Discrimination Index (2)</th>
<th>Item No.</th>
<th>Difficulty Index (4)</th>
<th>Discrimination Index (5)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0</td>
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Test items 1, 4, 5, 25 and 36 fall short of the standards set for the test but all other items are above the discrimination breakoff point of 20.

The difficulty index is a broad measure of the percent of the testees passing an item since a zero percent of success gives a difficulty index of zero and a hundred percent success gives a difficulty index of one hundred. The ratio is not 1:1 all the way...
through, though it is quite close to it. In this test a range of difficulty was desired and was achieved, the larger number of items (twenty-five) falling above the difficulty index of 50. The ceiling of the test was not reached by any tester as the largest number of items correct on any one paper was twenty-five.

**Determination of the reliability coefficient.**—The method of "rational equivalence" was used to determine the reliability coefficient of the Symbolic Unit-Test on equation-balancing. This method attempts to get an estimate of the reliability of a test by stressing intercorrelation of the item with the test as a whole, and bases its calculations on the proportion of correct and incorrect answers for the individual test items.

\[
\tau_{il} = \frac{n}{(n-1)} \times \frac{SD_i^2 - \hat{ip}q}{SD_t^2}
\]

By the use of the above formula the reliability was determined in terms of the difficulty and intercorrelation of the test items.

- \(\tau_{il}\) = reliability coefficient of the whole test
- \(n\) = number of test items
- \(SD_t\) = standard deviation of test items
- \(p\) = proportion of group answering test items correctly
- \(q\) = \((1-p)\) = proportion of group answering test items incorrectly

Application of the data taken from the first thousand papers corrected yielded the following results:

1/ Henry Garrett, op. cit. p. 325
2/ Henry Garrett, op. cit. p. 326
\[ r_{11} = \frac{56 \times 4.2^2 - 7.2}{4.2^3} \]

\[ r_{11} = 0.61 \]

Hence it may be said that the test as a whole possesses a reliability coefficient sufficiently high to warrant its use as a test for measuring achievement on the unit of chemical equation-balancing.

The correlation coefficient between pre-test and post-test, -

This was also determined, but on a sample of 100 papers chosen at random from those received. A second random sample of 100 papers \(^2\) was taken and a check run made on these. The figures obtained agreed.

The formula chosen for for this work was: \(^1\)

\[ r^2_{xy} = \frac{(N \sum xy - (\sum x)(\sum y))^2}{(N \sum x^2 - (\sum x)^2)(N \sum y^2 - (\sum y)^2)} \]

From the random 100 papers was as follows:

\[ \sum x \quad 1090 \]
\[ \sum y \quad 1706 \]
\[ \sum x^2 \quad 13182 \]
\[ \sum y^2 \quad 32514 \]
\[ \sum xy \quad 19952 \]
\[ N \quad 100 \]

\[ r^2_{xy} = \frac{(100 \times 19952 - (1090)(1706))^2}{(100 \times 13182 - (1090)^2)(100 \times 32514 - (1706)^2)} \]

\[ r = 0.51 \]

\(^1\) Henry Garrett, op. cit. p.142.

\(^2\) For treatment of the random hundred papers see the end of the chapter.
Summary of Symbolic Unit Test Data

1. Validity................................ face
2. Reliability\(^1\)...................... \(r_{11} = \cdot 61\)
3. Correlation coefficient\(^2\)......... \(r_{12} = \cdot 51\) (between pre- and post-test)
4. Discrimination indices ............. above 20
5. Difficulty indices ................... good range of items

\(^1\) Random hundred check on Kuder-Richardson reliability coefficient.

\[ r_{11} = \frac{n}{n-1} \times \frac{SD_T^2 - \sum pq}{SD_i} \]

\[ r_{11} = \frac{36}{35} \times \frac{(4.2)^2 - 751}{(4.2)^2} \]

\[ r_{11} = \cdot 614 \]

\(^2\) Random hundred check on Pearson coefficient of correlation between pre-tests and post-test scores of Symbolic Unit Test.

\[ r_{xy}^2 = \frac{(N \Sigma XY - (\Sigma X)(\Sigma Y))^2}{(N \Sigma X^2 - (\Sigma X)^2)(N \Sigma Y^2 - (\Sigma Y)^2)} \]

\[ r_{xy}^2 = \frac{100 \times 19789 - (1087)(1779)^2}{(100 \times 15169 - (1087)^2)(100 \times 15169 - (1779)^2)} \]

\[ r_{xy}^2 = \cdot 26 \]

\[ r_{xy} = \cdot 509 \]
CHAPTER VIII
SUMMARY AND CONCLUSIONS

Summary,—

1. A study of the comparison of inductive and deductive methods of teaching the secondary school unit on equation-balancing in chemistry was made during the 1957-1958 school year.

2. This phase of the study involved thirty-two schools and fifty-four classes. Classes averaged twenty-five students each and were located in cities across the nation.

3. Fourteen hundred chemistry students in the third and fourth years of high school (Grades XI and XII) were used in this study.

4. Half the classes (twenty-seven) were taught from deductive notes and laboratory directions supplied and designed by the writer. These were designated as control classes. The other half (twenty-seven) were designated as experimental classes and were taught by the inductive notes and laboratory guide sheets designed and supplied by the writer.

5. In all Type A schools the same teacher taught both the inductive and the deductive classes in the same school, thus limiting the variations in teacher motivation.

6. In all Type B schools inductive and deductive classes were
taught by different teachers.

7. All teachers following the deductive method used the same class lecture notes, unit outlines and demonstrations. All students used the same laboratory directions. All the material was designed by the writer, mimeographed, and sent to the teachers in the schools participating in the study.

8. All teachers of the inductive method used the same class lecture notes, unit outlines, laboratory notes and demonstrations. All the students used the same laboratory guide sheets. These materials were designed, mimeographed and distributed by the writer.

9. All teachers observed the same time schedule for the presentation of the unit. All administered the same test designed and provided by the writer. The test was administered before and after the unit was taught.

10. All tests were sent to the writer for scoring.

11. A mental ability test was given to all students participating in the study. 1/

12. Critical ratios were applied among the means-standard deviations of the intelligence test scores. These failed to reveal any significant variation in intelligence, hence, variation in chemistry achievement is possibly the result of variation in methods of teaching.

13. A t-test was applied among the differences in scores on

1/ The Quick Scoring Test of Mental Ability, op. cit.
pre- and post-tests on the Symbolic Unit-Test on equation-balancing in all Type A schools to determine if there was a significant difference in achievement produced by the inductive or by the deductive method of teaching. The data rejected the null hypothesis of equivalence and indicated the superiority of the inductive method.

14. A t-test was applied to the differences in scores of the pre-test and post-tests of the Symbolic Unit-Test in Type B schools and again the null hypothesis of equivalence in teaching methods was rejected. The inductive method was superior at the one percent level of confidence.

15. The application of the F-test to deviations of the pre-test scores failed to reveal any significant differences in the means of inductive classes of Type A or Type B schools. Deductive classes of Type A and Type B schools were also shown to be similar. Hence, all inductive classes were pooled for testing against all deductive classes.

16. A t-test was applied to the Symbolic Unit Test score differences between pre-tests and post-tests in all inductive and deductive classes. This test revealed the superiority of the inductive method at the one percent level of confidence.

17. By the single group method it was shown that there was a significant difference between pre-test and post-test scores of the upper quartile intelligence groups of inductive and deductive classes.

18. By the difference method a t-test was made on the post-test
scores of inductive and deductive groups of the upper quartile intelligence scores. This test indicated superiority of the inductive method at the one percent level of confidence.

19. A t-value was obtained for the post-test scores on the Symbolic Unit-Test for all inductive and deductive students of the lower quartile in intelligence as signified by intelligence test scores. A significant difference in favor of the inductive group was found to exist at the one percent level of confidence.

20. Item analyses of the Symbolic Unit-Test were made and the discrimination and difficulty indices of each item were determined.

21. Reliability and correlation coefficients were determined for the Symbolic Unit-Test.

Conclusions.

1. Inductive laboratory teaching produces higher achievement at the one percent level of confidence on the Symbolic Unit-Test on equation-balancing than does the deductive-descriptive laboratory teaching.

2. Students of both the upper and lower quartiles in intelligence as measured by an intelligence test were tested, and in both cases those inductively taught surpassed those deductively taught in their achievement on the Symbolic Unit-Test. This superiority was significant at the one percent level of confidence.

3. Students taught inductively are apparently able to grasp the

1/ Otis Quick Scoring Test of Mental Ability, op. cit.

2/ Ibid.
fundamental concepts of equation-balancing better than students taught by the deductive-descriptive method.

4. Consistent results in favor of the inductive method indicate that directions supplied by one teacher can be satisfactorily followed by other trained chemistry teachers.

General Conclusions from Phases I and II.--

The two phases of this study have yielded results which corroborate each other. In Phase I, twelve classes in three schools were used in an effort to determine the existence of significant differences in achievement in the field of high school chemistry under inductive and deductive methods of teaching. Results were calculated from scores on two standardized tests for high school chemistry. These results affirmed the superiority of the inductive method, which superiority was significant at the one percent level of confidence for average classes, for upper and lower quartile students on the basis of intelligence.

Phase II of the study made use of more than fifty classes in thirty-two schools and determined the effect of inductive and deductive teaching methods on one unit of high school chemistry—equation-balancing. Again, at the one percent level of confidence, the inductively taught classes were superior. This again included classes of average intelligence, of upper and lower quartile intelligence groups.

Anderson Chemistry Test, op. cit.
Cooperative Chemistry Test, op. cit.

Boston University
School of Education
Library
I. Introductory Work

A. The practical value of chemistry

1. New foods vary our diet
2. New medicines prolong our lives
3. New materials add beauty to our clothes
4. New alloys meet the exigencies of war times

B. Definition of chemistry - as the science which treats of matter and its changes

C. Explanation of the terms of the definition

1. Science - systematized knowledge
2. Matter - anything that has weight and occupies space
3. Elements - natural and transuranic
   a. Metals - lustrous, conductors
   b. Non-metals
4. Compounds - definition
5. Mixtures - definition - distinguish from compounds
6. Changes
   a. Physical
   b. Chemical - define and exemplify

7. Properties - characteristics by which a substance may be identified.
   a. Physical
      1. Color
      2. State - solid, liquid, gas
      3. Odor
      4. Hardness of solids
      5. Solubility
      6. Density as compared with air or water
b. Chemical
   1. Does or does not burn
   2. Does or does not support combustion
   3. Chemical activity

b. The metric system of weights and measurements

II. Oxygen

A. Occurrence
   1. Free in air
   2. Combined in water and in earth’s crust

B. Preparation
   1. Historical - Priestly heated mercuric oxide
   2. Commercial
      a. Electrolysis of water
      b. Liquefaction of air
   3. Laboratory - decomposition of potassium chlorate
      (catalytic action of I₂O₃)

C. Properties
   1. Physical: colorless, tasteless gas, slightly soluble in
      water; heavier than air
   2. Chemical: will not burn; will support combustion; active-
      combines with other elements forming oxides
      a. Explanation of terms
         1. Oxide
         2. Oxidation (rapid and slow)
         3. Combustion
         4. Spontaneous combustion
         5. Kindling temperature

D. Test - glowing splint glows more brightly or bursts into flame.

E. Uses
   1. Necessity to life
   2. Purification of water
   3. Oxygen tents
   4. Means of obtaining high temperature - torches
F. Allotropism "Varieties of the same element which differ in physical properties but have similar chemical properties are called allotropes."

1. Osme

   a. Preparation: electric discharge through moist O_2
   b. Properties: gas, pungent odor, nearly colorless, slightly soluble in water
   c. Uses: oxidizing agent, bleach, disinfectant, deodorizer

O. Vocabulary of new terms

allotropes oxidation kindling temperature
oxidation oxide rapid oxidation
oxidation catalytic agent oxidizing agent slow oxidation
combustion spontaneous combustion ordinary burning

III. Hydrogen

A. Occurrence

Combined in water, fuels, acids, sugars, living tissue

B. Preparation

1. Historical

   a. Contributions of Paracelsus and Boyle
   b. Henry Cavendish, 1766, first recognized it as a distinct substance
   c. Lavoisier named it hydrogen, i.e., water-former

2. Commercial

   Electrolysis of water

B. Laboratory

   a. Potassium, calcium, sodium on water
   b. Replacement in some dilute acids by certain metals

C. Properties

1. Physical

   a. Lightest gas known; greatest lifting power
   b. Diffuses through porous materials; explain osmosis
c. Adsorbed by platinum with liberation of heat - occlusion

2. Chemical

a. Combustible
b. Does not support combustion
c. Active - combines directly with sulfur and chlorine
d. Mixed with air it burns explosively
e. Reducing action on copper oxide

D. Test - Burning splint is extinguished and the gas burns with a pale blue flame.

E. Uses

1. Hydrogenation of oils
2. Balloons
3. Manufacture of water-gas
4. Synthesis of ammonia
5. Constituent of fuel gases
6. Refining of petroleum
7. Reducing agent

F. Isotopes - atoms of the same element having slightly different atomic weights

1. Protium - common hydrogen
2. Deuterium - heavy hydrogen - atomic weight 2; deuterium plus oxygen yields deuterium oxide, heavy water
3. Tritium - atomic weight 3

G. Vocabulary of new terms

diffusion adsorption occlusion
reduction displacement osmosis
hydrogenation electrolysis synthesis
reducing agent isotope water-gas

IV. WATER

A. Importance - relation of impure water to health of community
B. Composition - by weight and by volume
C. Analysis of water - by electrolysis
D. Properties
   1. Physical - including boiling and freezing points. It exists in all states: solid, liquid and gas.
   2. Chemical - treat heavy water here.

E. Methods of purifying water
   1. Boiling
   2. Filtration
   3. Distillation
   4. Oxidation
   5. Chlorination

F. Test - boiling and freezing points; also action in contact with anhydrous copper sulfate

G. Uses

II. Solutions
   1. Types - definition and method of preparing each type
      a. Unsaturated
      b. Saturated
      c. Super saturated
         1) Crystate
         2) Water of hydration (of crystallization)
         3) Hydrated and anhydrous conditions
      d. Molal solutions
      e. Normal solutions. These two types may be treated later in the course.
   2. Definition of terms: solution, solute, solvent, solubility, miscible, non-miscible
   3. Effect of dissolved substance on freezing and boiling point of solution
   4. Effect of temperature and pressure changes on the degree of solubility of solids; of gases

I. Suspensions and colloidal condition as opposed to true solutions - Brownian Movement.

J. Hydrogen peroxide
   1. Preparation - barium peroxide + sulfuric acid
2. Properties - thick, oily liquid, 1 1/2 times heavier than water, very unstable, decomposition products are water and nascent oxygen

3. Uses
   a. Bleaching by oxidation
   b. Insecticides
   c. Disinfectant

E. Vocabulary of new terms

- anhydrous
- anhydride
- boiling point
- melting point
- concentrated
- analysis
- volatile
- solvent
- solute
- saturated
- tincture

- crystalline
- crystallization
- dehydration
- efflorescent
- theory
- solubility
- suspension
- supersaturated
- condensation
- precipitate
- distillation

- evaporation
- freezing point
- hydrate
- miscible
- immiscible
- scientific law
- water of hydration
- water of crystallization
- dilute

V. LAWS, HYPOTHESES, AND THEORIES (Definition of the terms)

A. Law of definite proportions - illustrates by water
   Law of multiple proportions - illustrates by the oxides of hydrogen, nitrogen, and carbon

B. Law of conservation of matter - apply to equations

C. Dalton's atomic theory

D. Modern concepts of the atom - the electron theory

1. Units of structure
   a. Proton
   b. Neutron
   c. Electron
   d. Positron

   Characteristics of each as to mass and charge

2. Atomic number
   a. Nuclear charge
   b. Arrangement of electrons (hydrogen-calcium)
c. Isotopes (hydrogen and chlorine)

3. Explanation of periodicity in properties of elements in terms of atomic structure

A. Ionic valence
   a. Definition
   b. Methods of calculating
      (1) Memorize the valence of the most common elements and radicals
      (2) Deduce valence from a known formula
      (3) Determine valence from the atomic structure
   c. Relationship to inert gas structure
   d. Chemical reactions explained in terms of electrovalence
      (1) Combination
      (2) Displacement
      (3) Oxidation
      (4) Reduction
   e. Elements classified accordingly as
      (1) Metals
      (2) Non-metals
      (3) Amphoteric
      (4) Inert

5. Covalence – sharing of electrons in the formation of unionized molecules

B. Cases

1. Kinetic-molecular Theory
   a. Relation of temperature and speed of the molecules
   b. Differences between solids, liquids and gases from the kinetic-molecular viewpoint

2. Factors influencing reaction rate
   a. Concentration
   b. Temperature
   c. Catalysis
   d. State of division
3. Equilibrium - Mass Action Law, and the Principle of Le Chatelier as applied to gaseous reactions
4. Temperature, pressure and volume relations explained in terms of the kinetic theory
   a. Boyle's Law
5. Charles' Law - problems involving these laws
5. Gay-Lussac's Law
   Ratio of combining volumes and their products
6. Avogadro's Hypothesis
   Ratio of gram molecular volume, density and molecular weight
   1. Vapor density
   2. Molecular weight calculated from specific gravity referred to air, and from vapor density

F. Vocabulary of new terms

- molecule
- molecular weight
- symbol
- formula
- radical
- planetary electron
- binary compound
- ternary compound
- positive
- nucleus
- negative
- molecular formula
- atomic number
- electron
- proton
- shell
- orbit

VI. SIGNS, FORMULAS, AND EQUATIONS

A. Significance of the terms: symbol, formula, equation
B. Reading and writing formulas
C. Completing and balancing equations
D. Problems
   1. Determining percentage composition from the formula
   2. Determining simplest formula from percentage composition

VII. COPPER, A TYPICAL METAL

A. Occurrence
   Combined only
B. Preparation

1. Historical
   
   1744 - By electrolysis of melted NaOH

2. Commercial
   
   Electrolysis of fused NaCl

C. Properties

1. Physical
   
   a. Silver-gray color
   b. Lustrous
   c. Comparetively soft

2. Chemical
   
   a. Combustible
   b. Action with H₂O
      
      (1) Effect on litmus paper
      (2) Action with acids

D. Test - flame test yellow; examination of the spectroscope
   
   Spectrum analysis
   
   sodium  -------- yellow
   potassium  ------- violet
   lithium  ------- red
   barium  --------- green
   copper  -------- green
   strontium  ------- red
   calcium  --------- orange-red

E. Uses

F. Make a comparative study of potassium and sodium

VII. CHLORINE, A TYPICAL NON-METAL

A. Occurrence

Combined

B. Preparation

1. Historical - Karl Scheele, 1774, first prepared it by
   
   action of MnO₂ on H₂O; Davy, 1810, proved chlorine to be
   
   an element.
2. Commercial - electrolysis of NaCl solution

3. Laboratory
   a. Oxidation of HCl
   b. Action of NaCl with MnO₂ + H₂SO₄

C. Properties
   1. Physical
      a. Greenish-yellow gas
      b. Rancid odor
      c. Very soluble in water
      d. 2 1/2 times as dense as air
      e. Poisonous

2. Chemical
   a. Does not burn
   b. Does not support combustion; direct union reactions
   c. Action with water
   d. Action with non-metals
   e. Replacement series of non-metals

D. Uses
   1. Bleaching by oxidation
   2. Disinfecting
   3. Purifying water
   4. Preparation of organic compounds
   5. Poisonous gas in warfare

E. Test
   1. Free chlorine - odor and color; potassium-starch-iodide turns blue
   2. Chloride - silver nitrate causes a milky-white precipitate to form which is insoluble in KCl and soluble in NH₄OH.

F. Hydrogen chloride
   a. Preparation - sodium chloride + sulfuric acid
   b. Properties - colorless gas, irritating odor, very soluble in water
   c. Water solution is hydrochloric acid
   d. The gas is poisonous
d. The water solution obtained is a typical acid with
typical taste, action on litmus, reaction with
metals and neutralization of bases
e. The gas is 1/3 more dense than air
f. It is easily liquefied
g. It fumes in moist air.

II. ACIDS, BASES, AND SALTS

A. Acids

1. Composition - non-metallic element or radical plus
   hydrogen, often replaceable by a metal
2. Preparation - sulfuric acid plus a salt of the acid desired
3. Test - litmus turns red
4. Memorize formulas for the common acids
   a. HCl
   b. HNO₃
   c. H₂SO₃
   d. H₂SO₄
   e. H₃PO₄
5. Typical properties
   a. Contain hydrogen ions in solution
   b. The hydrogen is often replaceable by a metal
   c. They redden litmus
   d. Sour taste
6. Acid anhydride

B. Bases

1. Composition - metal with an hydroxyl radical
2. Preparation
   a. Active metal with water
   b. Union of oxide of an active metal with water
   c. Precipitation by mixing solution of the salt of
      a metal with a base
3. Test - turns litmus blue
4. Basic anhydride
C. Salts
Composition - metal plus an acid radical

X. PROBLEMS BASED ON THE CHEMICAL EQUATION
A. Volume-volume reactions
B. Weight-weight reactions
C. Weight-Volume reactions

XI. SULFUR AND THE SULFIDES
A. Occurrence - free and combined
B. Extraction - Frasch method
C. Allotropic forms
   1. Rhombic
   2. Prismatic
   3. Amorphous
D. Properties in general
   1. Amphoteric
   2. Combustible
   3. Moderately active
E. Uses
   1. Vulcanizing rubber
   2. Lime-sulfur sprays
F. Compounds
   1. Hydrogen sulfide
      a. Occurrence
      b. Preparation
      Ferrous sulfide + dilute HCl
   c. Properties
      (1) Physical
         (a) Odor
         (b) Solubility
         (c) Weight compared with air
      (2) Chemical
         (a) Combustion
         (b) Action with silver
(c) Formation of sulfides by double displacement

(d) Weak acid

d. Test for a sulfide

Dilute $\text{H}_2\text{SO}_4$ added to the unknown will turn lead acetate paper black if a sulfide is present.

2. Sulfur dioxide - sulfuric anhydride

a. Preparation

(1) Burning sulfur
(2) Action of acid on a sulfite

b. Properties

(1) Weight compared with air
(2) Reaction with $\text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$

c. Uses

(1) Disinfectant
(2) Bleaching
(3) Manufacture of $\text{H}_2\text{SO}_4$
(4) Refrigerating agent

3. Sulfurous acid

4. Sulfuric acid

a. Preparation - Contact Process (equations)

(1) Contact Process
(2) Lead Chamber Process

b. Properties

(1) Specific Gravity
(2) Boiling point
(3) Action of dilute acid with zinc
(4) Action of concentrated acid with copper
(5) Dehydrating action

c. Uses

(1) Preparation of other acids
(2) Manufacture of fertilizers
(3) Refining petroleum
(4) Storage batteries

d. Test for the sulfate ion
   Add barium chloride to the unknown. Very fine crystals are formed which are insoluble in water and dilute acids

C. Vocabulary of new terms

rhomboic       amorphous       prismatic
insecticide    vulcanization  dehydrating agent

XII. IONIZATION HYPOTHESIS - AVOGADRO

A. Electrolysis
   1. Definition
   2. Examples
      a. Water solutions of acids, bases, salts
      b. Saturated NaOH and NaCl

B. Non-electrolytes
   1. Definition
   2. Examples - solutions of
      a. Sugar
      b. Glycerine
      c. Alcohol

C. Conductivity explained in terms of ions

D. Contrast between ion and atom

E. Effect of ions on boiling point and freezing point of solutions

F. Definition of acids and bases in terms of ionization
   a. Acid salts
   b. Basic salts

G. Completion of ionic reactions
   1. Formation of insoluble solid or gas
   2. Formation of a slightly ionized product

H. Reversible reactions

I. Law of Mass Action

J. Neutralization explained by union of hydrogen and hydroxyl ions
K. Vocabulary of new terms

- acid
- salt
- base
- indicator
- ionic reaction
- alkaline
- neutralization
- titration

- ion
- ionization
- electrode
- hydrolysis
- dissociate
- equilibrium
- anode
- cathode

XIII. SODIUM COMPOUNDS (Make a comparative study of the potassium equivalents.)

A. Sodium hydroxide - caustic soda

1. Preparation - Oastner Process (electrolysis of brine)

2. Properties
   a. Strong base
   b. Corrosive
   c. Solvent for silk and wool
   d. Very soluble
   e. Deliquescent

3. Uses
   a. Soap making
   b. Bleaching solution
   c. Manufacture of wood pulp for paper
   d. Making rayon
   e. Refining petroleum

B. Sodium Peroxide

1. Preparation - heat chips of sodium in air free from CO₂

2. Uses
   a. Preparation of small amounts of O₂
   b. Preparation of H₂O₂
   c. Oxidizing and bleaching

C. Sodium Chloride

1. Occurrence
a. Sea water
b. Salt wells
c. Lakes
d. Salt lakes

2. Methods of extraction and purification

3. Uses

a. Manufacture of many chemicals
b. Freezing and packing ice cream
c. Preservative
d. Distigs of HI in gastric juice

3. Sodium Carbonate and Bicarbonate

1. Preparation - Solvay Process - memorize equations
2. Uses

4. Sodium Nitrate and Potassium Nitrate

1. Occurrence
2. Preparation
3. Properties
4. Uses

XIV. NITROGEN AND THE ATMOSPHERE

A. Nature of the atmosphere
1. Constituents
2. Proof that it is a mixture
   a. Composition varies
   b. No definite boiling point
   c. Can be separated by physical means
   d. No evidence of energy when air is synthesized
3. Uses of air
   a. It does work
   b. Supplies oxygen for breathing and combustion
   c. Enables us to hear sounds
   d. Furnishes earth with moisture
   e. Liquid air is a commercial source of Hz and Oz
4. Inert gases of the atmosphere
B. Nitrogen

1. Occurrence - free in air and combined in soil

2. Preparation
   a. Commercial - liquefaction of air
   b. Laboratory

3. Properties
   a. Physical
      (1) Lighter than air
      (2) Very slightly soluble in water
      (3) Colorless, odorless, tasteless
   b. Chemical
      (1) Relatively inactive
      (2) Does not burn
      (3) Does not support combustion
      (4) Active at high temperatures with H₂, O₂,
          CaO₂, and certain metals
      (5) Compounds are mostly unstable
      (6) Compounds frequently decompose explosively

4. Uses

XV. NITROGEN COMPOUNDS

A. Ammonia

1. Occurrence
   a. Atmosphere
   b. Soil

2. Importance
   a. Vital for fertilizers in peace
   b. Vital for explosives in war

3. Preparation
   a. Laboratory
      (1) Heat an ammonium salt with a base
(2) Nitride plus water

b. Commercial

(1) Labor Process
(2) Synamite Process
(3) Destinctive distillation of soft coal

4. Properties

a. Physical

(1) Colorless gas
(2) Argent odor
(3) Half as dense as air
(4) Very soluble in water

b. Chemical

(1) Relatively inert when dry
(2) Does not support combustion but will burn in part $O_2$
(3) Reducing agent
(4) Unites with some metals to form nitrides
(5) Reacts with $O_2$ in presence of catalyst yielding oxides of nitrogen
(6) Reacts with acids = ammonium salts
(7) Combines with $H_2O = NH_4OH$

5. Test for ammonium ion = ammonium salt plus strong base gives off the odor of ammonia. Red litmus will turn blue.

6. Uses

a. Refrigeration = because easily liquefied
b. Cleaning
c. Salicy Process
d. Satural Process
e. Manufacture of ammonium salts

3. Oxides of nitrogen

a. Nitric
b. Nitrous

2. Formulas

3. Preparation

4. Properties

5. Uses
C. Nitric Acid - aqua fortis

1. Preparation
   a. Laboratory - heating a mixture of NaNO₃ + H₂SO₄
   b. Commercial
      (1) Arc Process - Mix N₂ with an excess of O₂
          and subject it to a strong electric discharge
      (2) Ostwald Process - oxidation of NH₃

2. Properties
   a. Physical - colorless when pure
   b. Chemical - very active and highly ionized
      (1) Powerful oxidizing agent
      (2) Action with metals, particularly copper
      (3) Action with non-metals

3. Uses
   a. Manufacture of explosives
   b. Artificial silk
   c. Collodion, celluloid and plastics
   d. Dyes and drugs

4. Test for the nitrate ion
   Ferrous sulfate plus concentrated acid (sulfuric)
   forms a brown ring between the two layers.

5. Fixation of nitrogen
   a. Nitrogen cycle
   b. Artificial method

D. Vocabulary of new terms

- liquefy
- nitrifying bacteria
- nitrogen fixation
- relative humidity
- crop rotation
- anesthetic
- oxygen cycle
- photosynthesis
- nitrogen cycle
- aqua regia

XVI. THE HALOGENS

A. Fluorine, bromine, iodine compared with chlorine as to properties and activity
B. Preparation, properties and uses of bromine and iodine
C. Preparation of hydrogen fluoride - etching of glass
D. Test for bromide and iodide ions
E. Photography - preparation of the film, developing and fixing
F. Vocabulary of new terms

halogen antichlor activity series of non-metals
sedative sublimation

XVII. THE PERIODIC TABLE

Brief treatment giving variation of the elements illustrated in the halogen family and horizontally by argon through chlorine.

XVIII. CARBON

A. Occurrence
   1. Crust of the earth
   2. Living things
B. Importance
   700,000 compounds contain carbon
C. Allotropic forms
   1. Diamond
   2. Graphite
   3. Amorphous - hard and soft coal
D. Properties
   1. Physical
      Odorless, tasteless, relatively insoluble
   2. Chemical
      a. Inactive at ordinary temperatures
      b. Unites with O₂ at high temperature
      c. Will burn
      d. Reducing agent
      e. Unites with metals and non-metals at high temperatures to form carbides
D. Preparation of commercial forms
   a. Charcoal
b. Coke  
c. Boneblack  
d. Lampblack  
  Destructive distillation  

F. Compounds  

1. Carbon Dioxide  
  a. Occurrence  
    (1) In water and air  
    (2) Relation to plant and animal life  
  b. Preparation  
    Carbonates with acid  
  c. Properties  
    (1) Weight compared with air  
    (2) Solubility in water  
    (3) Non-supporter of combustion  
  d. Uses  
    (1) Beverages  
    (2) Fire extinguishers  
    (3) Leavening agent  
    (4) Fermentation  
    (5) Baking powders  
  e. Test - action in limewater  

2. Carbon monoxide  
  a. Preparation - concentrated H$_2$SO$_4$ + formic acid  
     or oxalic acid  
  b. Properties  
    (1) Physical  
      (a) Colorless  
      (b) Odorless gas  
      (a) Insoluble in water  
      (d) Same weight as air  
    (2) Chemical  
        (a) Combustibility  
        (b) Reducing agent
Ap oaloil.a

Action on red corpuscles

O odoruent of water gan and producer goa
d. Vocabulary of new terms
decolorizer bituminous coal
d. Activated charcoal adsorption

d. Calcium carbonate
occurrence in trouble-
limestones and calcites

2. Physical

(1) White solid
(2) Insoluble in water
(3) Stainable and stains

(1) Silvery white metal
(2) Harder than silver
(3) Heats alchohol with cold water releasing F
gas

B. Calcium carbonate
a. Physical

1. Occurrence - marble, limestone, calcite
2. Preparation - electrolysis of fused calcium chloride
3. Properties

B. Occurrence
1. Combined
2. Preparation - electrolysis of fused calcium chloride

a. Uses

Ca. Uses

b. Manufacture of CaO

m. Uses

1. Building
b. Manufacture of CaO

(1) Reacts with acids - CaO and a salt of the acid
(2) When heated, decomposes - CaO + Ca
(3) Hard waters - softening water
(4) Heats alcohol with cold water releasing F
gas

B. Calcium carbonate
a. Physical

1. Occurrence - marble, limestone, calcite
2. Preparation - electrolysis of fused calcium chloride
3. Properties

1. Calcium carbonate
2. Preparation - electrolysis of fused calcium chloride
3. Properties

a. Physical

1. Occurrence - marble, limestone, calcite
2. Preparation - electrolysis of fused calcium chloride
3. Properties

1. Calcium carbonate
2. Preparation - electrolysis of fused calcium chloride
3. Properties

a. Physical

1. Occurrence - marble, limestone, calcite
2. Preparation - electrolysis of fused calcium chloride
3. Properties

a. Physical
C. Calcium oxide

1. Preparation - heat CaCO₃ to obtain CaO + CO₂

2. Properties
   a. White solid
   b. High melting point
   c. Reacts with H₂O → Ca(OH)₂

3. Uses
   a. Manufacture of paper, rubber, building material
   b. Softening and purifying water

D. Calcium hydroxide

1. Preparation - H₂O + CaO

2. Properties
   a. White powder
   b. Slightly soluble in water
   c. Highly ionized in solution → strong base

3. Uses
   a. Preparation of CaO, NaOH, bleaching powder, whitewash, mortar, plaster-of-Paris
   b. Neutralizing soil acidity
   c. Limewater
   d. Water softening

4. Bleaching powder
   a. Preparation = Ca(OH)₂ + Cl₂ = chlorinated lime
   b. Uses = bleaching and disinfecting

E. Calcium sulfate

1. Gypsum - properties and uses

2. plaster-of-paris - properties and uses

3. Permanent hardness of water

F. Calcium phosphate

a. Occurrence - minerals, animal bones

b. Uses = fertilizers
3. Vocabulary of new terms
permanent hardness  
quicklime  
abrasive  
slake  
stalactite  
stalagmite

XX. MINERALS
A. Important mineral substances not previously discussed.
Processes, formulas, and equations not required.
1. Garnalite
a. Constituents - K2O, MgCl2-6 H2O
b. Use as source of commercial potassium compounds
to be used as fertilizers
2. Rock Phosphate - Ca3(PO4)2
a. Uses
   (1) Fertilizer
   (2) Source of phosphorus
   (5) Phosphoric acid
   (4) Monocalcium phosphate
   (5) Disodium phosphate
3. Galena
a. Formula - PbS
b. Use - principal source of lead

XXI. METALS
A. Definition - elements forming positive ions in solution and
whose soluble oxides are basic
B. Characteristic properties - malleability, ductility, conductivity,
color, luster, crystalline structure
C. Relative activity of metals - memorize the electrochemical series
D. General principles of metallurgy
   1. In case of native metals such as gold
   2. Reduction of oxides, e.g. iron oxide reduced by Carbon
   3. Fouling followed by reduction in the case of sulfides
      and carbonates
4. Electrolysis in the case of aluminum oxide

E. Refining of metals
   1. wrought iron and steel from pig iron
   2. Electrolysis for refining copper

F. Electric furnace products
   a. Electric furnaces
      1. Arc
      2. Resistance
   2. Preparation, properties, uses of the following:
      a. Calcium carbide
      b. Carborundum
      c. Calcium cyanamide
      d. Phosphorus
      e. Carbon disulfide

G. Identification of the ions
   1. Aluminum
   2. Copper
   3. Iron
   4. Zinc

II. Alloys
   1. Sterling silver
   2. Coin silver
   3. Brass
   4. Bronzes
   5. Solder
   6. Carbon steels
   7. Chrome
   8. Tungsten
   9. Molybdenum

XXII. ORGANIC CHEMISTRY

A. Hydrocarbons — definition— principal series with general
   formulas and examples in each series
   1. Methane \( \text{CH}_4 \)
a. Graphic formula
b. Occurrence - natural oil fields and marshy regions
c. Use - fuel gas

2. Acetylene $C_2H_2$
   a. Graphic formula
   b. Preparation -- water + CaC$_2$ → 2 Ca(OH)$_2$ + $C_2H_2$
c. Uses - illuminating and welding

3. Benzene $C_6H_6$
   a. Graphic formula
   b. Source - destructive distillation of coal
c. Use - solvent

B. Mixture of hydrocarbons - petroleum products
   1. Gasoline
   2. Kerosene
   3. Lubricating oils
   4. Vaseline
   5. Paraffin

Fractional distillation and cracking

C. Compounds related to hydrocarbons - substitution products of methane
   1. Chloroform, iodoform, carbon tetrachloride
      a. Learn the chemical name, formula, chief use of
         the above
   2. Alcohols
      a. Definition - an organic compound containing the
         hydroxyl group
      b. Kinds
         (1) Wood alcohol
         (2) Grain alcohol
         (3) Glycerine
   3. Organic acids - contain the carboxyl group -COOH
      a. Acetic acid - CH$_3$COOH
      b. Source - oxidation of alcohol
      c. Properties - weak acid; peculiar odor
d. Uses - making vinegar, cellulose, dying silk and wool

D. Foods

1. Classification
   a. Proteins
   b. Fats - soap making and by-products
   c. Carbohydrates
      (1) Definition
      (2) Kinds
         (a) Starch
         (b) Cellulose
         (c) Sugars
APPENDIX B
INDUCTIVE SYLLABUS

UNIT I - STRUCTURE OF MATTER

I. Conditions of matter
   A. Gases and vapors
      1. Description in terms of kinetic-molecular theory
      2. Metric units of measurement of volumes
   B. Liquids
      1. Description in terms of kinetic-molecular theory
      2. Units of measurement of volume and weight
   C. Solids
      1. Description in terms of the kinetic-molecular theory
      2. Metric units of measurement of volume, weight and size
      3. Crystalline matter
         a. Lattice of polar compounds
            (1) Properties dependent on structure
            (2) Coordinate bonds
         b. Lattice of non-polar compounds
            (1) Properties dependent on structure
            (2) Covalent bonds
         c. Ionic compounds
            (1) Proton donors .... acids
            (2) Proton acceptors .... bases
            (3) Salts
   4. Amorphous matter
   5. Allotropes
   6. Density

II. Classes of matter
   A. Elements and their symbols
   B. Compounds
C. Mixtures

III. Dalton's Atomic Theory
   A. Original statements
   B. Posts still retained

IV. Modern Atomic Theory
   A. Structure - Bohr Theory
   B. Atomic number
   C. Atomic weight

       1. Relative to oxygen
       2. Isotopes

D. Valence

   1. Metals ----- electron donors
   2. Non-metals ----- electron acceptors
   3. Inert elements ----- complete orbit
   4. Ionic valence
   5. Covalence
   6. Coordinate valence

V. Vocabulary for Unit I

   absorption                      diffusion                      liter
   allotrope                      dipole                         mass
   alloy                           distillate                    number
   amorphous                       efflorescent                  mass
   anhydrous                       electron                      metal
   atom                            electronnegeutive              mixture
   atomic number                   electropositive                molecule
   atomic weight                   element                        monoprotic
   binary compound                 energy level                   neutron
   boiling point                   evaporation                    non-metal
   chemical change                 filtrate                       non-polar compound
   combustion                     fluid                          nucleus
   compound                       fractional distillation         orbital electron
   condenser                       fractional crystallization      proton
   covalent bond                   gas                            radical
   covalent linkage                gram                          solid
   crystal                         hydrate                        sublimation
   crystallization                hygrosopic                     substrate
   dehydrating agent              inorganic                       substance
   dehydrated                      ion                            symbol
   deliquescent                    isotope                         ternary compound
   density                         kilogram                       valence electron
   dissilator                     kindling temperature           vapor
UNIT II - PERIODIC TABLE

I. Historical development
A. Early attempts at classification
B. Moseley
C. Modern Periodic Table - isotopes; fractional atomic weights
D. Advantages of such a classification

II. Structure of the Table
A. Groups - same number of electrons - vertical rows
   1. Valence to hydrogen  O = 1 --4--1
   2. Valence to oxygen  O =---------- 7
   3. Activity
      a. Metals - increase with increase of atom size
      b. Non-metals - decrease with increase of atom size
   4. Subgroups
      a. Valence electrons the same, but family characteristics differ.
      b. Activity decreases with increase in weight
B. Series - horizontal rows - ring structure of the kernel of the atom the same as that of the inert gas at the beginning of the series.
   1. Character
      Inert gas --- metals --- amphoteric --- non-metals
   2. Activity
      Metals --- amphoteric --- non-metals
   3. Transition elements - caused by variation in capacities of inner shells.
      Fe 2-8-14-2
      Co 2-8-15-2
      Ni 2-8-16-2
   4. First series - Hydrogen's unique place in the table
   5. Short series - second and third - eight elements
   6. Long series - fourth - eighteen elements
      fifth - eighteen elements
      sixth - thirty-two (rare earths)
      seventh - thirty-two by prediction (transuranics)
7. Analysis of several groups and series

III. Vocabulary for Unit II

<table>
<thead>
<tr>
<th>Word</th>
<th>Combustion</th>
<th>Covalence</th>
<th>Lanthanide series</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acid</td>
<td>Covalence</td>
<td>Covalence</td>
<td>Malleability</td>
</tr>
<tr>
<td>Activity</td>
<td>Coordinate covalence</td>
<td>Deuterium</td>
<td>Metal</td>
</tr>
<tr>
<td>Alkali metal</td>
<td>Ductility</td>
<td>Non-metal</td>
<td>Non-metal</td>
</tr>
<tr>
<td>Alkaline earth metal</td>
<td>Effervescence</td>
<td>Oxidation</td>
<td>Oxidation</td>
</tr>
<tr>
<td>Alloy</td>
<td>Fixation of nitrogen</td>
<td>Flame</td>
<td>Ozone</td>
</tr>
<tr>
<td>Anhydride</td>
<td>Free elements</td>
<td>Inert</td>
<td>Pneumatic trough</td>
</tr>
<tr>
<td>Aqua regia</td>
<td>Inert</td>
<td>Rare earth elements</td>
<td>Properties</td>
</tr>
<tr>
<td>Atom</td>
<td>Inert</td>
<td>Rare earth elements</td>
<td>Rare earth element</td>
</tr>
<tr>
<td>Atomic number</td>
<td>Inert</td>
<td>Rf</td>
<td>Salt</td>
</tr>
<tr>
<td>Atomic weight</td>
<td>Inert</td>
<td>Rf</td>
<td></td>
</tr>
<tr>
<td>Base</td>
<td>Inert</td>
<td>Rf</td>
<td></td>
</tr>
<tr>
<td>Catalyst</td>
<td>Inert</td>
<td>Rf</td>
<td></td>
</tr>
</tbody>
</table>

IV. Laboratory for Unit II

A. Group VI, Oxygen and Sulfur

1. Preparation and properties of oxygen and a study of the types of oxidation.
2. Study of the properties of sulfuric acid; dilute and concentrated.
3. Preparation of hydrogen sulfide; determination of its use as an analytical reagent.
4. Preparation and study of the properties of sulfur dioxide; test for the sulfate ion and the sulfite ion

B. Group V, Nitrogen

Preparation and study of the properties of ammonia; test for ammonia and the ammonium ion

C. Group IV, Carbon

Preparation and study of the properties of carbon dioxide; test for carbon dioxide and for the carbonate ion.

D. Group II, Alkaline Earth Metals

1. Flame tests for calcium and barium
2. Hard waters and how to soften them

E. Group I, Hydrogen and the Alkali Metals

1. Flame tests for sodium and potassium
2. Preparation and study of the properties of hydrogen; test for hydrogen gas and the hydrogen ion
3. Action of sodium on water; test for hydroxyl ion
UNIT III - SIMPLE EQUATIONS AND THE LAWS INVOLVED

I. Word reactions
   A. State the experimental facts accurately
   B. Write reactants on the left side of the arrow; products on the right.

II. Symbolized equations from word equations - write correct formulas
    for all reactants and products.

III. Balanced equations
    A. Law of Conservation of Matter
    B. Make the total number of atoms on left of arrow equal total
       number on right, using smallest common denominator as
       balancing number.

IV. Meaning of equations
    A. The reactants and products of the equation
    B. The composition of the molecules that react
    C. The number of each kind of molecule taking part in the reaction
    D. The weights of the reactants and the products
    E. The volume of the gaseous substances

V. Classification of reactions
    A. Direct combination or synthesis
    B. Simple decomposition or analysis
    C. Simple or single replacement or substitution
    D. Double replacement, double decomposition or metathesis

VI. Reactions that go to completion
    A. By the formation of a slightly ionized substance
    B. By the formation of a gas
    C. By the formation of an insoluble substance; rules of solubility

VII. Equilibrium
    A. Law of Mass Action
    B. Le Chatelier's Principle
    C. Common Ion Effect

VIII. Energy changes
    A. Exothermic and endothermic reactions
    B. Heat of formation
II. Vocabulary for Unit III

- analysis
- combining number
- combining weights
- decomposition
- displacement
- double decomposition
- double replacement
- dynamics
- energy
- endothermic reaction
- exothermic reaction
- equation
- equilibrium
- equivalent
- formula
- gram atomic weight
- hypothesis
- metathesis
- mole
- neoncent
- oxide
- precipitate
- products
- reactant
- radical
- reaction
- reagent
- reversible
- solubility
- single replacement
- synthesis
- unstable compound
- volatile

III. Laboratory for Unit III

A. LA Series
B. Synthetic and Analysis
C. Single and double replacement
D. Double decomposition reactions that go to completion
E. Four ways of preparing a salt

UNIT IV - PROBLEMS BASED ON LAWS OF CHEMICAL COMPOSITION

I. Dalton's original theory and present modifications

II. Law of definite composition or proportion

A. Atomic weights
   1. Relative on basis of oxygen taken as 16
   2. Average of weights in the case of isotopes
   3. Relation of atomic weights to atomic number and neutrons

B. Formula of a compound
   1. Relative weight of a molecule - mole and g.m.w.

C. Calculation of molecular weights

D. Calculation of percentage composition

Supplement with exercises illustrating Law of Multiple Proportions

E. Calculation of simplest formula
   1. With molecular weights given
2. Without molecular weight given
3. Simplest formula indicates simplest possible ratio between whole atoms of the compound
4. Supplementary exercises utilizing valence variations

III. Law of Conservation of Mass

A. Expressed in the balanced equation
B. Quantitative equations involving weight
C. Quantitative calculations involving weight and volume of ideal gas
   1. Gram molecular volume = 22.4 liters
   2. Standard Temperature and Pressure = 273°C and 760mm

IV. Law of Combining Volumes and Avogadro’s Law

Calculations involving volume and the use of formulas and equations

V. Vocabulary for Unit IV

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>avogram</td>
<td>gravimetric</td>
</tr>
<tr>
<td>density</td>
<td>ideal gas</td>
</tr>
<tr>
<td>gram atomic weight</td>
<td>mass</td>
</tr>
<tr>
<td>gram molecular volume</td>
<td>mole</td>
</tr>
<tr>
<td>gram molecular weight</td>
<td>kinetic-molecular theory</td>
</tr>
</tbody>
</table>

UNIT V - SOLUTIONS, SUSPENSIONS, EMULSIONS, COLLOIDS

I. Solutions

A. Solvent and solute
B. Types of solutions
   1. Gas in gas .............. review gas laws
   2. Gas in liquid ..... ... Henry's Law; temperature effects
   3. Liquid in liquid .... effect of temperature, agitation, grinding, nature of the solute, etc.
   4. Saturated, unsaturated and supersaturated solutions
   5. Raoul's Law
      1. Freezing point depression and boiling point elevation
      2. Vapor pressure depression
E. Osmosis

II. Suspensions and emulsions
   A. Characteristics
   B. Size of particles
   C. Separation of constituents

III. Colloids
   A. Characteristics
      1. Size of particle
      2. Brownian movement
      3. Tyndall Effect
      4. Osmophoresis
      5. Dialysis
   B. Preparation
      1. Peptization
      2. Dispersion
      3. Condensation
      4. Mechanical disintegration
      5. Emulsification
   C. Terminology
      1. Dispersion medium
      2. Dispersed phase
      3. Emulsoid .... liquid in liquid
      4. Suspensoid .... solid in liquid
      5. Dispersoid ..... general

IV. Vocabulary for Unit V

activated charcoal      electron microscope      millimicron
adsorption             gel                          saturated
Brownian movement      homogeneous
coagulation            heterogeneous
Cottrell separator     heterogeneous
colloidal state        immiscible
dilute                 lyophobic
dispersion             lyophilic
peptization            membrane
emulsion               miscible

Tyndall Effect
UNIT VI - ACIDS, BASES, AND SALTS

I. Acids
   A. Definition .... proton donor or electron acceptor
   B. Characteristics
      1. Sour taste
      2. Turns blue litmus red
      3. Turns phenolphthalein colorless
      4. Contains hydrogen ions that can be replaced by metals
         to form salts
      5. Neutralizes bases
   C. Preparation
      1. Anhydride and water
      2. Salts and water
      3. Salts and sulfuric acid
      4. Direct combination (for binary acids)
   D. Hydrochloric acid, a typical acid
   E. Naming of acids ....
      1. Binary
      2. Ternary
   F. Ionization
      1. Arrhenius theory
      2. Debye-Hückel Theory
      3. Hydronium ion
      4. pH scale and indicators

II. Bases
   A. Definition ..... proton acceptor or electron donor
   B. Characteristics
C. Preparation
   1. Oxides and water
   2. Salt and a strong base
   3. Active metal and water
D. Sodium hydroxide, a typical base
E. Ionization and reaction toward indicators
F. Naming bases

III. Salts
   A. Definition
   B. Characteristics
   C. Preparation
      1. Neutralization
      2. Acid on metal
      3. Metal and non-metal in direct union
      4. Acid on metallic oxide
      5. Acid on salt of a more volatile acid
      6. Reaction between two salts
      7. Reaction between basic and acidic oxides
D. Kinds
   1. Normal salt
   2. Acid salt
   3. Basic salt
E. Hydrolysis
F. Naming of salts

IV. Vocabulary for unit VI

<table>
<thead>
<tr>
<th>acid</th>
<th>cation</th>
<th>litmus</th>
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<tr>
<td>acid anhydride</td>
<td>dissociation</td>
<td>monobasic acid</td>
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<td>acid salt</td>
<td>dibasic</td>
<td>monoprotic acid</td>
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<tr>
<td>acidic oxide</td>
<td>dissociation</td>
<td>neutralization</td>
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<td>anhydride</td>
<td>electrode</td>
<td>normal salt</td>
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<tr>
<td>anion</td>
<td>electrolyte</td>
<td>pH</td>
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<td>anode</td>
<td>electrolysis</td>
<td>phenolphthalein</td>
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<td>aqua regia</td>
<td>hydrolysis</td>
<td>salt</td>
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<tr>
<td>base</td>
<td>hydroxide</td>
<td>strong electrolyte</td>
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<tr>
<td>basic oxide</td>
<td>indicator</td>
<td>ternary compound</td>
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<tr>
<td>binary compound</td>
<td>ion</td>
<td>titration</td>
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<tr>
<td>cathode</td>
<td>ionization</td>
<td>tribasic</td>
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</table>
V. Laboratory for Unit VI

1. Preparation of soluble and insoluble salts
2. Properties of acids and bases
3. Hydrolysis
4. Acid and basic anhydrides
5. Neutralization

UNIT VII - PROBLEMS INVOLVING GASES

I. Kinetic-molecular theory
   A. Notion of molecules
   B. Elasticity
   C. Temperature effects

II. Charles' Law applied to an ideal gas
   A. Conversion of Fahrenheit to Centigrade and Absolute scales
   B. Calculation of gas volumes with relation to temperature changes

III. Boyle's Law applied to an ideal gas
     Variations in pressure and volume of gases

IV. Concentration and standard solutions
   A. Equations involving water solutions
   B. Molar solutions
   C. Normal solutions
      1. Gram equivalent weight per liter
      2. Molecular weight divided by number of electrons lost
      3. Normality x volume = normality x volume

V. Vocabulary for Unit VII

<table>
<thead>
<tr>
<th>Term</th>
<th>Definition</th>
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<tbody>
<tr>
<td>absolute temperature</td>
<td>barometer</td>
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<td>centigrade scale</td>
<td>standard solution</td>
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<tr>
<td>Fahrenheit scale</td>
<td>qualitative</td>
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<tr>
<td>Kelvin scale</td>
<td>quantitative</td>
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<tr>
<td>Celsius scale</td>
<td>density</td>
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<td>molar</td>
<td>molecular weight</td>
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<tr>
<td>molal</td>
<td>avogram</td>
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<tr>
<td>mole</td>
<td>mass</td>
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<tr>
<td>normal solution</td>
<td>gram atomic weight</td>
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<tr>
<td>equivalent weight</td>
<td>gram molecular weight</td>
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<tr>
<td>standard conditions</td>
<td>ideal gas</td>
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<tr>
<td>S.T.P.</td>
<td>gram molecular volume</td>
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UNIT VII - OXIDATION - REDUCTION

I. Definition and determination of oxidation number or state
   A. In ionic compounds
   B. In covalent compounds

II. Definitions and explanations of
   A. Oxidation
   B. Reduction
   C. Oxidizing agent
   D. Reducing agent

III. All single replacement reactions are redox reactions.

IV. All direct combination reactions are redox reactions.

V. Balancing redox reactions
   A. Assign oxidation numbers
   B. Mark changes per atom in oxidation state
   C. State number of atoms needed to use electrons in change
   D. State number of molecules needed to supply these atoms
   E. Balance H⁺ and OH⁻ with water
   F. Write complete equation

VI. Relative oxidizing and reducing power of substances
   A. Review electromotive series
   B. Metals work as reducing agents on lower ions
   C. Ions work as oxidizing agents on higher metals

VII. Reactions of powerful and important oxidizing agents
   A. Concentrated sulfuric acid
   B. Concentrated and dilute nitric acid
   C. Manganese dioxide and permanganate
   D. Halogen preparation reactions
   E. Oxidation of ferrous to ferric ions

VIII. Reactions of powerful and important reducing agents
   A. Hydrogen on copper oxide
   B. Carbon monoxide and carbon as commercial reducers in metal refining

IX. Determination of the equivalent weight of an oxidizing agent and of
    a reducing agent.
X. Vocabulary for Unit VIII

- electron
- equivalent weight
- mineral
- ore
- oxidation
- oxidizing agent
- oxidation number
- oxidation state
- reduction
- reducing agent

XI. Laboratory for Unit VIII

1. Reducing action of hydrogen
2. Reducing action of carbon (metals from ores)
3. Chlorine -- preparation and properties
   a. Test for chlorine
   b. Test for the chloride ion
4. Bromine -- preparation and properties
   a. Test for bromine
   b. Test for the bromide ion
5. Iodine -- preparation and properties
   a. Test for iodine
   b. Test for the iodide ion
6. Nitric acid -- preparation and properties
   a. Test for the nitrate ion
   b. Test for nitric oxide
   c. Test for nitrogen dioxide
7. Ferrous-ferric changes and the identification of Fe²⁺ and Fe³⁺
8. Sulfuric acid -- properties; test for the sulfate ion

XII. Ion-electron method of balancing redox equations (optional)

A. Advantages: Valence state is never mentioned. No assumption of fractional valence is required. In most aqueous solutions the ions react and not the molecules, as are indicated in molecular equations.

B. Necessary concepts: In each redox reaction, we have an oxidation "half cell" and a corresponding reduction "half-cell". These must be understood, and by taking any combination one may balance the reaction involved.

C. Oxidation "half-cells" and their reduction products (not balanced)
1. \( \text{NO}_2^- + H^+ \rightarrow \text{NO}_2 + H_2O \) (conc.)
2. \( \text{NO}_3^- + H^+ \rightarrow \text{NO} + H_2O \) (dil.)
3. \( \text{MnO}_4^- + H^+ \rightarrow \text{Mn}^{++} + H_2O \)
4. \( \text{Cr}_2\text{O}_7^{2-} + H^+ \rightarrow \text{Cr}^{+++} + H_2O \)
5. \( \text{ClO}_3^- + H^+ \rightarrow \text{Cl}^- + H_2O \)
6. \( \text{FeO}_4^{2-} + H^+ \rightarrow \text{Fe}^{+++} + H_2O \)
7. \( \text{MnO}_4^- + H^+ \rightarrow \text{Mn}^{+++} + H_2O \)
8. \( \text{Cl}_2 + H^+ \rightarrow \text{Cl}^- + \)

D. Reduction "half-cells" and their oxidation products (not balanced)
1. \( S^{--} \rightarrow S \text{ or } \text{SO}_2 \text{ or } \text{SO}_3 \)
2. Metals \( \rightarrow \) metal ions (e.g. Cu \( \rightarrow \text{Cu}^{++} \))
3. \( \text{Cl}^- \rightarrow \text{Cl}_2 \)
4. \( \text{Br}^- \rightarrow \text{Br}_2 \)
5. \( I^- \rightarrow I_2 \)
6. \( \text{AsO}_3^{--} \rightarrow \text{AsO}_4^{---} \)
7. \( \text{I}^- \rightarrow I_2 \)
8. \( \text{SO}_2 \rightarrow \text{SO}_3 \)

E. Reactions illustrating the method: According to our concept, we can take any item from paragraph C and react it with one from paragraph D and obtain an equation.

As an example, let us take C-3, and D-4. Thus we have an ionic skeleton: \( \text{MnO}_4^- + \text{Br}^- + H^+ \rightarrow \text{Mn}^{+++} + \text{Br}_2 + H_2O \) (not balanced)

Next we set up the oxidation "half-cell" and balance it electronically, adding enough hydrogen ions to "burn" off all the oxygen atoms in the MnO\(_4^-\), and producing water. Adding sufficient electrons to the proper side of the equation to give electrical neutrality, we arrive at the following:
\[ 5\text{E}^- + \text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{+++} + 4\text{H}_2O \] (note: there is a net of 2+ on each side)

Next we do the same for the reduction "half-cell":
\[ 2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^- \]

We note that two electrons are liberated on the right of the second "half-cell" while five are required on the left of the initial "half-cell" reaction. We must make those evolved from the second
equal to those consumed by the first. Obviously, we multiply the first by two and the second by three. After doing this the electrons cancel and by addition of the equations we arrive at the balanced, correct, ionic equation:

\[ 2\text{MnO}_4^- + 10 \text{Br}^- + 16 \text{H}^+ \rightarrow 5 \text{Br}_2 + 2 \text{Mn}^{2+} + 8 \text{H}_2\text{O} \]

Note that ions, charges and molecules are all balanced and we have not mentioned valence:

\[ 2 \text{MnO}_4^- + 16 \text{HBr} \rightarrow 2 \text{KBr} + 2 \text{MnBr}_2 + 5 \text{Br}_2 + 8 \text{H}_2\text{O} \]

XIV. Equations to be balanced:

1. \[ \text{PbO}_2 + \text{H}_2\text{S} \rightarrow \text{PbS} + \text{SO}_2 + \text{H}_2\text{O} \]
2. \[ \text{K}_2\text{Cr}_2\text{O}_7 + \text{HCl} \rightarrow \text{KCl} + \text{CrCl}_3 + \text{H}_2\text{O} + \text{Cl}_2 \]
3. \[ \text{Na}_2\text{CrO}_4 + \text{NaClO} + \text{NaOH} \rightarrow \text{NaCl} + \text{Na}_2\text{CrO}_4 + \text{H}_2\text{O} \]
4. \[ \text{Cu} + \text{HNO}_3 (\text{conc}) \rightarrow \text{NO}_2 + \text{Cu(NO}_3)_2 + \text{H}_2\text{O} \]
5. \[ \text{Cu} + \text{HNO}_3 (\text{dil}) \rightarrow \text{NO} + \text{Cu(NO}_3)_2 + \text{H}_2\text{O} \]
6. \[ \text{KMnO}_4 + \text{H}_2\text{O} + \text{KI} \rightarrow \text{MnO}_2 + \text{KOH} + \text{KIO}_3 \]
7. \[ \text{KClO}_3 \rightarrow \text{KClO}_4 + \text{KCl} \]
8. \[ \text{KMnO}_4 + \text{FeSO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{MnSO}_4 + \text{Fe}_2(\text{SO}_4)_3 + \text{H}_2\text{O} \]
9. \[ \text{Cu} + \text{H}_2\text{SO}_4 \text{ heated} \rightarrow \text{CuSO}_4 + \text{SO}_2 + \text{H}_2\text{O} \]
10. \[ \text{KMnO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{MnSO}_4 + \text{H}_2\text{SO}_4 + \text{H}_2\text{O} \]
11. Aluminum + hydrochloric acid
12. \[ \text{Ag} + \text{HNO}_3 \rightarrow \text{AgNO}_3 + \text{H}_2\text{O} + \text{NO} \]
13. \[ \text{CrI}_3 + \text{KOH} + \text{Cl}_2 \rightarrow \text{K}_2\text{CrO}_4 + \text{KIO}_4 + \text{KCl} + \text{H}_2\text{O} \]

UNIT IX - RADIOACTIVITY

I. Review of terminology of atomic structure

II. Radioactivity and atomic structure
   development of properties of radium on this basis

III. Radiations
   A. Alpha rays
   B. Beta rays
   C. Gamma rays

IV. Disintegration
   A. Transmutation
   B. Half-life
V. Laws of conservation in the light of Einstein's equivalence equation
VI. Nuclear binding energy
VII. Artificial radioactivity
   A. Transmutation
   B. Exploding nuclei
VIII. Types of nuclear reactions
   A. Alpha particle bombardment (Rutherford)
   B. Neutron bombardment (Fermi)
      1. Fission
      2. Fusion
      3. U-235
   C. Deuteron bombardment
IX. Chain reaction
   A. Atomic pile
   B. Atomic bomb
   C. Hydrogen bomb
X. Radioisotopes
   A. Tracers
   B. Other uses
XI. Vocabulary for Unit IX

alpha particle  atomic particle  beta particle  binding energy  cathode ray  chain reaction
artificial radioactivity  atomic pile  beta particle  binding energy  cathode ray  chain reaction
fission  fusion  gamma ray  half-life  heavy hydrogen  heavy water
neutron  nucleus  nuclear bombardment  positron  radionuclide  slow neutron
prefix  prefix  prefix  prefix  prefix  prefix
other uses
XII. Laboratory work for Unit IX (not related to the unit contents)
   1. Identification of silver, lead, and mercurous mercury ions
   2. Rocking powder analysis
   3. Identification of ions in unknowns
**APPENDIX G**

**INDUCTIVE LABORATORY GUIDE SHEETS**

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<td>Percentage of water in copper sulfate</td>
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<td>Determination of the molecular volume of oxygen</td>
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<td>Types of chemical reactions</td>
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<td>Preparation and properties of acids and bases</td>
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<td>Ionization</td>
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<td>Normal and molar solutions</td>
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<td>25</td>
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<td>Preparation and properties of hydrogen sulfide</td>
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<td>Preparation and properties of chlorine</td>
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<td>Preparation and properties of bromine</td>
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<tr>
<td>Preparation and properties of iodine</td>
<td>34</td>
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</table>
GENERAL OUTLINE REPORT FORM

A. Occurrence
   1. Place and state
   2. Method of extraction (if any)

B. Preparation
   1. Laboratory (description and equations)
   2. Commercial (description and equations)

C. Properties
   1. Physical
      a. Density
      b. Color
      c. Odor
      d. Taste
      e. Solubility
      f. Change of state
      g. Special (allotropes, isotopes, etc.)
   2. Chemical
      a. Combustibility
      b. Stability
      c. Reactions with elements
      d. Reactions with compounds
      e. Special (oxidizing agent, acid, etc.)
      f. Test for identification

D. Structure
   1. Atomic diagram
   2. Molecular diagram
   3. Ionic diagram

Placement in Periodic Table
Metal or non-metal?

E. Uses

F. Other important equations involving this substance

G. Diagram of laboratory preparation
Measure four objects in the laboratory and record the data in the following table:

<table>
<thead>
<tr>
<th>Object</th>
<th>Length (cm)</th>
<th>Length (in.)</th>
<th>Width (cm)</th>
<th>Width (in.)</th>
<th>Thickness (cm)</th>
<th>Thickness (in.)</th>
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</tbody>
</table>

Express the metric measurements in millimeters and meters.

Measure your test tubes and record the linear dimensions. Fill the tubes with water and pour the water into a graduated cylinder to determine the volume of the tubes. Practice estimating 1, 2, 3, 4, 5 ml, amounts in your test tubes. Also estimate 10, 25, 50 ml quantities in your beakers and flasks so that you will not have to measure the substances you use in experimentation.

If you are going to use a balance this year, accustom yourself to its operation now.

Practice conversion of metric units into other metric units and into English units.

THE BUNSEN BURNER

1. Principle of the burner—Note the construction of the burner. Be sure you notice the air vent and the small pin-hole opening for the gas. Practice lighting the burner and adjusting the flame. Determine what is the hottest part of the flame, the reducing flame and the oxidizing flame.
I. To a four-inch test tube, add about a gram (spatulaful) of ferric salt; add 5 ml. \( \text{K}_2\text{Cr}_2\text{O}_7 \) solution and 2 drops of conc. \( \text{H}_2\text{SO}_4 \). Keep this tube corked.

II. Evaporation. -- Place a few drops of the mixture prepared in I in a crucible or evaporating dish and heat GENTLY until the liquid almost disappears.

III. Centrifugation. -- Stir the original mixture and quickly pipet off about 10 drops into a centrifuge tube half filled with water. Fill another tube to the same extent with water and place both tubes in the centrifuge. Let the machine run for about half a minute then remove the tube.

IV. Precipitation. -- To 2 ml. of \( \text{H}_2\text{O} \) in a centrifuge tube, add about 8 drops of the centrifugate obtained in III and add 2 ml. \( \text{Pb(CH}_3\text{O})_2 \). Balance this tube with one of water and centrifuge.

V. Distillation. -- Pour half the remainder of the original mixture into a 4-inch test tube fitted with a one-hole stopper and delivery tube leading to a small test tube. Place this second test tube in a beaker of ice. Heat the mixture until about 1 ml. of distillate is collected. Test it with litmus paper. (If a Bunsen burner is used, heat very gently holding the burner in the hand and moving it over the tube. Less caution is needed if a micro burner is used.)

VI. Filtration. -- Fold a filter paper in quarters and slip it into a funnel. Stir the remainder of the original mixture and quickly pour it or pipet it onto the filter. Catch the filtrate in a small test tube or on a watch glass.
Preliminary Experiments

The following three preparations are used in the collection of gases. Facility in gas collection will be a great asset later in the course.

Gas Set-up "A"— This is used for gases insoluble in water. Place a small piece of moist zinc in a 4-inch test tube fitted with a stopper and conventional delivery tube, as in the diagram. Fill a pneumatic trough with water and immerse 5 vials or 4-inch test tubes. Add about 3 ml. dil. H₂SO₄ to the zinc and insert the delivery tube so that the end is beneath the surface of the water in the trough. Slip a collecting vial over the end of the delivery tube and allow it to fill with gas. When the tube is full of gas (empty of water) remove it and replace it with another vial. Inserting tubes full of gas, cover the end of the tube with the thumb, or insert a cork while holding the tube under water.

Gas Set-up "B"— This is used for gases lighter than air and soluble in water. To a generator tube (4-inch test tube) add about a gram of M₄O₁ and an equal amount of Ca(OH)₂. Attach an oblique angle delivery tube and heat the generator gently. Collect the gas in vials by the upward displacement of air. A piece of moist litmus placed at the end of the collecting vial will turn color when the vial is filled.

Gas Set-up "C"— This is used for gases heavier than air and soluble in water. To about a gram of salt in a generator
tube, add 3 ml. dil. H₂SO₄ and attach the acute angle delivery tube. Collect the gas by the downward displacement of air.

SUGGESTIONS TO STUDENTS

Perform all experiments while holding equipment at arm's length. Equipment is small and simple, no vessel being over 25 ml capacity. When a test is to be made on a liquid, the solution to be tested is placed in a 5 ml test tube, filling it up to a depth of about 1 ml. To this is added, dropwise, the reagent.

To determine odor, waft the fumes toward the nose with the cupped hand.

Transfer of solids.— Never pour directly from a bottle to a flask or test tube but transfer the solid to a small piece of creased, glazed paper, and, using this as a slide, slip the solid into the test tube.

Transfer of liquids.— Turn the palm of the hand upward and grasp the stopper of the reagent bottle with the index finger and third finger, and remove it. Holding the stopper between the two aforementioned fingers, grasp the bottle and pour the liquid into the tube. Small amounts of liquids may be transferred by means of a dropper or pipet.

Measuring volume.— Graduated cylinders, burettes, and pipets are used to measure the volume of liquids. In reading the level of the liquid in these tubes, keep the eye on a level with the bottom of the meniscus and read the line upon which it rests.

Testing with indicator paper.— Transfer a drop of the solution
to be tested to the test paper by means of a stirring rod.

Heating materials.— This is always done in Pyrex, thin-walled glassware or porcelain. Thick-walled bottles and not heat resistant. When heating test tubes, hold them at an angle pointing away from you.

Filtration.— Support the funnel in a ring stand and tough the stem to the wall of the receiving vessel. Fold the paper in halves along the diameter and again in halves. Open the quarter nearest you and place the cone in a moist funnel. Ease it in until it fits snugly.

Centrifugation.— The material to be centrifuged must be placed in a tube which fits the centrifuge and must be counterbalanced by a similar tube containing water. A half-minute of turning generally suffices to settle any precipitate.

Testing solubility.— Of gases: Invert a tube of the gas in a beaker of water. After some minutes, water will rise in the tube if the gas is dissolving. Of solids: Any evidence of dissolving action like the evolution of a gas or the disappearance of the solid is taken as indicative of solubility.

Testing combustion.— A burning splint inserted in a gas will either burn or go out, and the gas will ignite or will not. If the gas burns, efforts should be made to identify the products of combustion.

Density.— This may be inferred from the type of apparatus used in the collection of a gas, or may be calculated ahead of time if the apparatus set-up in unknown. Relative density is compared with that of air taken as 1.0."
GLASS MANIPULATION

I. Discussion.—In bending or drawing glass tubing, correct results are obtained only when the glass is rotated in the top part of the flame and is heated sufficiently long to make the glass bend very easily and seem as if it will break or fall into two parts. Also, in bending tubing, it should not be twisted, but rotated during heating.

II. Manipulation.—Using a file or an ampoule cutter, make a scratch (do not saw) about 6 inches from the end of a piece of glass tubing. Pick up the tubing, placing the thumbs opposite and on both sides of the scratch, and curl the fingers over the tubing. Break the tubing as you would break a piece of wood. (Fig. 1)

![Fig. 1](image)

After the tubing has been broken, adjust a wing tip on your burner and regulate the burner so as to obtain a luminous flame; then rotate the tubing in the flame. (Fig. 2) During the rotation, heat the same area of the tubing until it becomes soft and readily bendable. At this stage, remove the tubing from the flame and bend it to an obtuse angle as in the diagram. (Fig. 3) After the proper bend has been obtained, set the tube on an asbestos mat to cool. When cool, file the ends of the tubing until all rough pieces are chipped away and then heat both ends in the flame to obtain a smooth fire-polished finish. Repeat using 10 and 12 inch pieces of tubing and make the figures illustrated below.
Making a medicine dropper. — Cut a piece of glass tubing (4 mm outside bore) into 6 inch length. Without using a wing-tip, heat the tubing in the center and rotate until the tubing is very soft. Remove from the flame and draw the tubing apart until a long, narrow capillary portion is obtained. (Fig. 6) Place the tubing on an asbestos mat to cool. Be careful never to lay hot tubing on the desk top or contact with the cold surface will cause the glass to crack. After the tubing has cooled, cut out the capillary section into two pieces each 3 inches long. Now take one of these sections and heat the wide end until it becomes very soft; remove from the flame and press down on an asbestos mat. This last operation gives a flared end over which the rubber bulb is fitted. Again allow the glass to cool; finally polish the small end by heating it a very short time in the flame; attach the rubber bulb. Repeat with the other section of tubing. Test out those pipets by drawing up water from a test tube. There should be no leakage from the pipet. Present these, along with your glass bend, to your instructor for approval.

Making a spatula. — Cut a glass rod 6 inches in length. Heat one end until it is very soft. Remove from the flame and immediately press the heated end vertically on an asbestos mat or charcoal block. This gives a rounded flat end. Now heat the other end and press it at a 45° angle against a block or mat. This gives the other end of the rod as shown in the diagram. (Fig. 7)
Making a wash bottle.— If a 250 ml. Florence flask is used, make one obtuse angle bend in a piece of glass tubing 5 inches long. Then take a foot of tubing and make an acute angle supplementary to the obtuse angle already constructed. (A ruler can lie flat along tubes A and B in the drawing.) Fit this latter bend through a two-hole stopper which fits the flask in use and gauge the place where the lower bend must be made. Make a dropping pipet about an inch long and attach it with rubber tubing to the acute angle B. Fire polish all rough ends and assemble as in the diagram. (Fig. 8)

PARTMENTS, COMPOUNDS, AND MIXTURES

I. Collect the necessary material.—

- salt
- sand
- beaker
- filter paper
- funnel
- ring stand
- stirring rod
- mercuric oxide
- sulfur
- iron filings
- mortar and pestle
- magnesium ribbon
- copper metal
- (centrifuge tubes)

II. Perform the experiment.— Mix a pinch of salt with an equal portion of sand and add water. Filter or centrifuge. By testing at each step, determine the path the salt follows.

- Heat a pinch of MgO and test the combustion properties of the gas.

- Mix equal quantities of sulfur and iron filings and separate by the use of a magnet.

- See if the application of heat changes any of the properties of these substances.
I. Thoroughly test a small quantity of salt. Note its color, taste, and texture. Then grind it, dissolve it,蒸发ate off the water and examine the residue.

II. Determine the physical properties of a strip of magnesium ribbon before and after ignition.

III. Determine the physical properties of copper wire before and after heating.

IV. Can you clearly distinguish between physical and chemical changes and between mixtures and compounds? Make a list of ten of each; if any substance presents a difficulty, see if you can apply some laboratory tests to determine exactly what sort of substance it is.

DISTILLATION

I. Collect the materials—

- test tube generator and delivery tube
- NiCl₂
- Phenolphthalein
- CuSO₄ crystals
- beaker of ice
- small collecting tubes

II. Perform the experiment.— Dissolve a crystal of CuSO₄ in water. Set up equipment to distill this mixture. Have it approved before using.

Add a drop of ammonia to fresh water in the generator tube and distill the mixture. Device a method of testing the distillate for the presence of ammonia.
CRYSTALS AND CRYSTALLIZATION

I. Collect the necessary materials—

\begin{align*}
\text{CuSO}_4 & \quad \text{Cobalt chloride paper} \\
\text{CaCl}_2 \text{ anhydrous} & \quad \text{K}_2\text{Cr}_2\text{O}_7 \text{ crystals} \\
\text{NaOH} & \quad \text{glass plates} \\
\text{NaCl} & \quad \text{Na}_2\text{SO}_4 \\
\text{Alum} & \quad \text{Hypo} \\
\text{NH}_4\text{Cl} & \quad \text{CuSO}_4 \text{ anhydrous}
\end{align*}

II. Perform the experiment.—

See if any of the crystals listed contain water by heating them in a crucible and testing the fumes for \( \text{H}_2\text{O} \). If you can remove the water from any, see if you can replace it. See if any salts take on water from the atmosphere. Try to make a few large crystals of \( \text{CuSO}_4 \) or \( \text{K}_2\text{Cr}_2\text{O}_7 \).

SULFUR — ITS ALLOTROPIC FORMS

I. Collect the necessary materials.—

\begin{align*}
\text{Flowers of sulfur} & \quad \text{Filter paper} \\
\text{Ball sulfur} & \quad \text{Funnel} \\
\text{Carbon disulfide} & \quad \text{Beaker of water}
\end{align*}

II. Perform the experiment.—

By carefully heating sulfur, observe the various changes it undergoes. Prepare a sample of each of the allotropic forms. Remember the melting points and work accordingly. Submit your plans to your instructor before using.

III. Report in chart form the following information:

Form, viscosity, color, melting point, method of preparation, solubility
OXYGEN - A STUDY OF ITS PREPARATION AND PROPERTIES

I. Collect the necessary materials:
- KClO₃ .......... oxygen source
- MnO₂ .......... catalyst
- Sb powder
- Charcoal, powder
- Sulfur, powder
- Generator tube and collecting tubes
- Pneumatic trough
- Wooden splints
- Clamp stand and clamps and burner

II. How much KClO₃ would be needed to generate 40 ml. of O₂? Calculate the density of the gas.

III. Perform the experiment.— Remembering that O₂ is slightly soluble in water and that it is evolved when KClO₃ is heated, devise a safe experiment to collect 40 ml. of pure O₂ gas. Use apparatus of preliminary experiment "A" in gas collection. With the 4 vials of gas collected, test combustibility, reaction with powdered charcoal and with powdered sulfur. Observe the physical properties of O₂.

IV. Test the remaining solid in the generator to prove it is a chloride. Look up standard tests for the detection of O₂ and apply to some of the gas.

V. Fill in your outline.— Look up material on ozone, the allotropic form of O₂, and make an outline on its preparation and properties.

Any tests you have not performed to your satisfaction may be repeated. If there is no evidence for some part of the outline, test by your own method. Submit finished outline to your instructor for checking.
HYDROGEN - A STUDY OF ITS PREPARATION AND PROPERTIES

I. How much acid would you need to react completely with one gram of zinc? How much hydrogen would be generated? Calculate the density of H₂.

II. Collect the necessary materials—

- dil. H₂SO₄  .......... source of hydrogen
- Zn, moosy  .......... releasing agent
generator and collecting vials
- wooden splints
- pneumatic trough
- clamp, stand, burner, etc.

III. Perform the experiment— Hydrogen is slightly soluble in water.
Collect and test some of the gas. Make special efforts to check its relative density.

IV. Test for hydrogen. — Consult your text for the standard test for hydrogen gas and apply it. Clearly distinguish between the test for the gas and the test for the hydrogen ion.

V. Test the materials remaining in the generator to prove the course of the reaction.

VI. Fill in your outline.— Remember that hydrogen is a reducing agent.

ELECTROMOTIVE SERIES AND HYDROGEN PREPARATION

I. Collect the necessary materials—

<table>
<thead>
<tr>
<th>Na metal</th>
<th>Ca metal</th>
<th>HCl, dilute and concentrated</th>
</tr>
</thead>
<tbody>
<tr>
<td>Co metal</td>
<td>Zn metal</td>
<td>H₂SO₄ &quot; &quot; &quot;</td>
</tr>
<tr>
<td>Mg metal</td>
<td>Fe metal</td>
<td>HCl &quot; &quot; &quot;</td>
</tr>
<tr>
<td>Al metal</td>
<td>Mg metal</td>
<td></td>
</tr>
<tr>
<td>Sn metal</td>
<td>Co metal</td>
<td></td>
</tr>
<tr>
<td>Pb metal</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
II. Perform the experiment.— Test the reaction of sodium and calcium with water using a minute quantity of the metal in a covered beaker. Test the products for identity.

Using the suggested acids and metals, determine the relative ease with which a metal will replace hydrogen from an acid. Compare the acids as well as the metals. Using the HCl as the standard releasing agent, compare the metal. Test all products.

III. Report the experiment.— Arrange the metals in the descending order of reactivity; arrange the acids in the descending order of reactivity; balance equations for every reaction that occurred.

Write out a more complete EMF series from your text and check off in red those elements which release H₂ from H₂O; in blue, those which release H₂ from an acid; in green those which do not release H₂ at all.

DETERMINATION OF THE VALENCE OF MAGNESIUM

I. Collect the necessary materials.—

- a weighed piece of magnesium ribbon (not over 0.1 gr.)
- one 4-ounce bottle filled with water + HCl
- pan of water
- graduated cylinder
- glass plate
- barometer and thermometer

II. Perform the experiment.— React the magnesium with the water–acid mixture in the bottle by inverting the bottle in a pan of water. Level off the resulting gas and measure the volume it occupies. Record the temperature and pressure.
III. Report the results in the following form:

Weight of Mg taken ...................................................
Volume of hydrogen collected ........................................
Temperature of the water ............................................
Barometric pressure in mm ...........................................
Volume of Mg corrected for STP .....................................
Weight of hydrogen evolved .........................................
Wt. of hydrogen evolved by 1 atomic weight of Mg ..........
Valence of magnesium ................................................

PERCENTAGE OF COPPER IN COPPER OXIDE

I. Collect the necessary materials:--

2 clamp stands
buret pan balance and weights
obtuse angle delivery tube, straight tube
calcium chloride
copper oxide
zinc
sulfuric acid, dilute
generator tube
burner

II. Perform the experiment:-- Weigh the copper oxide before and after the experiment to determine the weight of oxygen lost by the process of reduction. This oxygen can be removed by passing fresh dry hydrogen gas over hot copper oxide. Set up your equipment and have your procedure approved before starting.

III. Report your experiment in the following form:
III. Report your results in the following fashion:

- weight of tube + copper oxide
- weight of tube alone
- weight of copper oxide
- weight of tube + copper after heating
- weight of copper
- percentage of copper
- theoretical percentage of copper
- percent of error

DETERMINATION OF THE MOLECULAR VOLUME OF OXYGEN
and
THE AMOUNT OF OXYGEN IN KClO₃

I. Collect the materials needed.

- material as needed in the oxygen preparation
- 4-ounce collecting bottle
- graduated cylinder
- thermometer
- barometer
- 1/216 weighing balance and weights

II. Perform the experiment.

Set up the apparatus for oxygen preparation, but weigh the
KClO₃ added. (Between 0.2 and 0.25 gr.) Measure the amount of oxygen
formed and collected in a 4-ounce bottle. Record the temperature
and pressure.

III. Report the results of the experiment in the following form.
weight of test tube generator + KClO₃ ..........................
weight of tube alone ...........................................
weight of KClO₃ ..............................................
weight of tube + KClO₃ + MnO₂ ..............................
weight of tube after heating .................................
weight of oxygen evolved ..................................
volume of oxygen in collecting bottle .....................
temperature of water .........................................
barometric pressure in mm .................................
volume of oxygen at STP ....................................
volume occupied by one gram of oxygen ...................
molecular volume of oxygen .................................
percentage of oxygen in KClO₃ ..............................
theoretical molecular volume of oxygen ..................
percent error ..................................................
theoretical percentage of oxygen in KClO₃ .............
percent error ..................................................

**TYPES OF REACTIONS**

There are four major types of chemical reactions: analysis, synthesis, single replacement, and double replacement.

I. Collect the necessary materials.-- You are asked to perform two experiments of each of the types listed above. Identify all products of reaction.

II. Perform the experiment.-- Report the methods you applied and give complete balanced equations for every reaction used.
REACTIONS THAT GO TO COMPLETION

I. Collect the necessary materials.—You are to prepare for and perform two reactions that go to completion by the formation of a gas; two, because a precipitate is formed; and two because a slightly ionized substance results.

II. Perform the experiments.—

III. Report the experiments.—Write ionic equations for each of these reactions and indicate which ions are removed from the field of action and how.

ACIDS AND BASES

Determine the common properties of acids and bases, both chemical and physical. List those in parallel columns so that you can see the similarities and differences between acids and bases.

I. General methods of preparing an acid.—

1. Salt of the acid desired + sulfuric acid
2. Acid anhydride + water
3. Direct combination (for binary acids)

II. General methods of preparing a base.—

1. Salt of the base desired + sodium hydroxide
2. Basic anhydride + water
3. Active metal + water

III. Prepare an acid and a base by one of the above methods and test the properties of the product.

Could you start with a piece of magnesium ribbon and prepare a
base? Try it.

Could you make an acid from a piece of sulfur? If it is possible to carry out such an experiment, do it.

Report this day's laboratory work by indicating the reactions that occurred and by balancing those reactions in order.

**IONIZATION**

I. Collect the necessary materials. —

- Conductivity apparatus
- Fifteen common acids, bases and salts
- Three organic substances

II. Perform the experiment. — Carry out the conductivity measurements and report the degree of conductivity on the following chart.

III. Write ionization reactions for those substances which ionize.

IV. Report strong electrolytes in order of increasing activity.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Brightness of Light</th>
<th>Type of Electrolyte</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td></td>
<td></td>
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<tr>
<td>2.</td>
<td></td>
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<tr>
<td>3.</td>
<td></td>
<td></td>
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<tr>
<td>4.</td>
<td></td>
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<tr>
<td>5.</td>
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<tr>
<td>6.</td>
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<td>7.</td>
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<td>8.</td>
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<td>9.</td>
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<tr>
<td>10.</td>
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<tr>
<td>11.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>12.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
SALTS

I. General methods of preparation.-- (For soluble salts)
   1. Metal + an acid
   2. Metallic oxide + an acid
   3. Carbonate + an acid
   4. Neutralization of a base with an acid
   5. Direct combination (for binary salts)

II. General methods of preparation.-- (For insoluble salts)
   1. Salt + an acid
   2. Salt + another salt
   3. Oxide + an acid

Review your tables of solubility and prepare two soluble and two insoluble salts.

SALTS SOLUBLE AND INSOLUBLE

I. General rules.-- To prepare pure soluble salts from soluble salts, the intermediate stage must be insoluble; and conversely, to prepare insoluble salts from insoluble salts, the intermediate stage must be soluble. To go from a soluble salt to an insoluble salt and vice versa, direct transformation is possible.

II. Perform the experiment.-- Prepare NaCl by at least four different methods; prepare BaSO₄ also by four methods. Give detailed explanation, with equations, of every step you use. Report this experiment by a flow sheet of equations.
HESOVLYSIS

I. Collect the necessary materials.

- 15 salts selected at random from the shelf
- small test tubes
- litmus paper

II. Perform the experiment. — Test a solution of each salt for
acidity by the use of litmus paper and report your results in short
form. Write balanced ionic equations for all reactions.

<table>
<thead>
<tr>
<th>Salt Formula</th>
<th>Litmus Reaction</th>
<th>Acid Formed</th>
<th>Weak or Strong</th>
<th>Base Formed</th>
<th>Weak or Strong</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2.</td>
<td></td>
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<td></td>
<td></td>
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<tr>
<td>3, etc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

TYPES OF SOLUTIONS

I. Collect the necessary materials.

- NH₄OH
- C₆H₅OH
- NaCl

II. Perform the experiment. — From the above chemicals prepare a
solution of:

- a. a gas in a liquid
- b. a liquid in a liquid
- c. a solid in a liquid

III. Working with a solid in liquid solution, show the effects of
stirring, of solvent, and of concentration on the rate of solubility.

IV. Prepare an unsaturated, a saturated, and a supersaturated solution
of sodium acetate or sodium thiosulfate and prove which type of solution you have by devising a simple test that will differentiate in all cases in which this determination must be made.

SOLUTIONS — NORMAL AND MOLAR

I. Collect the necessary materials. —
   burette and pipet
   indicator (liquid form)
   HCl
   KOH

II. Remembering the definitions of normal and molar solutions, calculate the amount of NaOH necessary to prepare 250 ml. of 0.01N NaOH, and then make up this solution.

   Using this known solution, calculate the strength of an unknown acid which you will obtain from the instructor.

III. In turning in your report, be sure to include all calculations. Report not only the normality and molarity of the unknown acid, but also the number of grams of solid per liter.

COLLOIDS

I. Collect the necessary materials. —
   FeCl₃
   gelatine
   starch
   beakers and test tubes
   burner, clamp and stand

II. Perform the experiment. —
   Devise an experiment to produce colloidal starch from a coarse
starch from a coarse starch suspension. Devise an experiment to form colloidal ferric hydroxide. Can you think of a way to utilize gelatine as a protective colloid? If so, try it out.

Prepare a summary of the physical properties of solutions, suspensions and colloids in parallel columns so that the corresponding similarities and differences are apparent at a glance. To do this you may have to prepare a few solutions and suspensions and test them. If so, go ahead.

If apparatus is available, try the Tyndall Effect on solutions, suspensions, and colloids. If there is no satisfactory vessel and light source, perhaps you could make one out of school time and demonstrate this point on colloid chemistry to the class.

HYDROCHLORIC CHLORIDE - ITS PREPARATION AND PROPERTIES

How much acid do you need to prepare 50 ml. of HCl gas? Calculate the density of the gas.

I. Collect the necessary materials.—

HgO & .......... source of hydrogen ions
HCl .......... source of chloride ions
generator tubes
4 small collecting vials
delivery tube, clamp stand and clamps

II. Perform the experiment.— You should prepare this acid by the reaction between a salt and an acid. Why does this reaction go to completion? Prepare both gaseous hydrogen chloride and liquid hydrochloric acid. Make out report outline sheets for both.
III. Test for HCl. -- Consult your textbook for the standard test for the presence and identification of HCl. Also find the test for the chloride ion and apply it to the solution of HCl you prepared.

IV. Test the material remaining in the generator. -- Prove the course of the reaction by proving the products of the reaction.

AMMONIA - A STUDY OF ITS PREPARATION AND PROPERTIES

I. Collect the necessary materials. --

- NH₄Cl ........ source of NH₃
- Ca(OH)₂ .......... releasing agent
generator tube and collecting vials
- litmus paper
- (NH₄)₂SO₄ ...... another source of NH₃
clamp stands, clamp and burner

II. Calculate the density of NH₃.

III. Perform the experiment. -- NH₃ is extremely soluble in water. Set up apparatus for the collection and testing of 25 ml. of the gas. Do not fail to observe its reaction with gaseous HCl.

IV. Test for NH₃. -- In your textbook, find the standard test for the identification of NH₃ and apply it to a sample of the gas. Apply the test for the identification of the NH₄⁺ ion also.

CARBON DIOXIDE - ITS PREPARATION AND PROPERTIES

I. Collect the necessary materials. --

- marble chips .......... source of CO₂
dilute HCl ........... releasing agent
generator tube and collecting vials
I. Collect the necessary materials.--

FeS .......... source of the sulfide ions
HCl .......... source of the hydrogen ions
generator and collecting vials
delivery tube
lead acetate paper
solutions of Cd^{++}, Sb^{++}, Cu^{++}, Pb^{++}, Zn^{++}, As^{+++}

II. Calculate the density of hydrogen sulfide. How much FeS is needed to produce 25 ml. H_{2}S?

III. Perform the experiment.-- Hydrogen sulfide is a poisonous gas that is soluble in water. Devise a safe method for the collection of 25 ml. of it. After the vials of gas are collected, allow the delivery tube to dip into a vial of water and bubble the gas through the water for a few minutes.

Observe the effect of the H_{2}S solution on solutions of heavy metallic ions, and also, for a more pronounced effect, bubble H_{2}S gas directly through solutions of those ions.

IV. Test for H_{2}S.-- Look up the standard test for H_{2}S and also for the sulfide ion and apply them to your solutions. Apply them also to a sample of the gas.

SULFUR DIOXIDE - ITS PREPARATION AND PROPERTIES

I. Collect the necessary materials.--

Na_{2}SO_{3} or NaHSO_{3} ............ source of SO_{2}
dilute H_{2}SO_{4} ................. releasing agent
generator and delivery tube
clamp stand and clamps
wooden splints
burner
beaker of water
KClO₄ solution
litmus paper

II. Calculate the density of SO₃. What volume of SO₃ is formed by the complete combustion of 3.2 grams of sulfur? What weight of SO₃ gas is generated by the complete combustion of 2.5 grams of sulfur? What weight of SO₂ gas is formed by the decomposition of 2.5 grams of Na₂SO₄? What volume would it occupy?

III. Perform the experiment. — This gas is quite soluble in water; use apparatus for preliminary experiment B. Devise an experiment for collecting a small quantity of SO₂. Test for all properties you can.

IV. Test the materials remaining in the generator. — Trace the path of the reaction positively by determining the identity of the substances remaining in the generator. Also, look up the standard test for the identification of SO₂ and of a sulfite and apply them.

OXIDIZING AGENTS

I. Collect the materials necessary. —

K₂MnO₄
KMnO₄
H₂SO₄ conc.
HNO₃ conc. and dilute
any other oxidizing agents you wish
substances to be oxidized

II. Perform the experiment. — Perform a few oxidation reactions and prove the products are what your equations indicate them to be.
Carefully record all tests made; write their equations. Be sure to test the different concentrations of KMnO₄ on common metals like Mn and Cu.

III. Write redox equations for all procedures and clearly indicate the electron changes, the oxidizing agent, the reducing agent, the substance oxidized, and the substance reduced.

REDUCING AGENTS

I. Collect the necessary materials. —

- materials for the preparation of hydrogen
- CuO, wire form
- carbon
- other substances to be reduced
- test tubes, burner, clamps, etc.

II. Perform the experiment. — Devise an experiment for reducing an oxide to its metallic constituent and prove you have the desired product. Consult your text for the commercial reduction of ores. See if you can reproduce any of these on a laboratory scale.

III. Write redox reactions for all these reactions and list the important industrial processes that are based on oxidation-reduction.

FERRIC = PERIODIC REACTIONS

I. Collect the necessary materials. —

- powdered iron
- potassium ferrioxalate
- HzO
- potassium thiocyanate
- potassium ferrocyanide
- NaOH
- KMnO₄
II. Perform the experiment.-- Dissolve some of the iron in acid and form solutions of ferric and ferrous ions.

Test the reactions of these solutions with NaOH, potassium ferrocyanide and potassium ferricyanide, and potassium thiocyanate.

Oxidize and reduce ferrous and ferric ions and test for the completeness of your oxidations and reductions.

III. Report redox reactions in equation form, along with descriptions of all color changes that occurred.

CHLORINE - ITS FORMATION AND PROPERTIES

I. Collect the necessary materials.--

\[
\begin{align*}
\text{NaCl} & \quad \text{oxydizing agent} \\
\text{HCl} & \quad \text{source of the chlorine generator and delivery tube} \\
\text{collecting vials and stoppers} & \\
\text{clamp stand, clamps, burner, etc.} & \\
\end{align*}
\]

II. Perform the experiment.-- Device an experiment for the safe and swift collection of four vials of chlorine gas remembering that the gas is extremely poisonous and soluble in water. Perform enough tests on the gas to complete your outline, and in addition, endeavor to ascertain the bleaching power of the gas and of its water solution.

III. Test for chlorine.-- Look up the standard test for chlorine and for the chloride ion. Identify all materials left in the generator.

BROMIDE - ITS FORMATION AND PROPERTIES

I. Collect the necessary materials.--
KBr .......... source of Br₂
MnO₄ ...... oxidizing agent
H₂SO₄ ...... oxidizing agent
generator and delivery tube
collecting vial
beaker of ice

II. Perform the experiment. -- WARNING ***** Bromine is extremely poisonous and the liquid and vapor forms are tissue destroyers.
If any Br₂ gets on you, flood the area with water and rub well with glycerine.

De devise an experiment for the rapid and safe collection of a few drops of pure bromine by the condensation of liquid bromine from the vapors given off when KBr is oxidized.

III. Consult your text for the activity series of non-metals and verify the placement of all the elements in it.

IV. Test for a bromide and for bromine. -- Apply the standard test for the identification of bromine and a bromide. Clearly distinguish between the test for the molecule and for the ion. All equations should be balanced by the oxidation-reduction method.

IODINE - ITS PREPARATION AND PROPERTIES

I. Collect the necessary materials.--
KI .......... source of I₂
NaO₂ ...... oxidizing agent
H₂SO₄ ...... oxidizing agent
small crucible and watch glass to fit

II. Perform the experiment. -- The equipment used is quite different
from that commonly employed in experimental preparations because iodine is a crystal and has been discovered to have the peculiar property of sublimation. Devise an experiment to produce some tiny pure crystals of iodine by oxidizing KI in a glass covered crucible.

Perform the tests necessary for the completion of your outline. Be sure to test the solubility of iodine in alcohol, water plus potassium iodide, and carbon tetrachloride.

III. Test for iodine. -- Look up the standard test for iodine and the iodide ion and apply them to samples of each. Identify the materials remaining in the generator. Write redox reactions for all reactions that are mentioned. Include in your final report a second outline on hydrogen iodide which you may write with the aid of your textbook. Hl will not be prepared in the laboratory.
## Tests and Equations to Be Known

<table>
<thead>
<tr>
<th>Substance</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia</td>
<td>odor, if base is added</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>lime water</td>
</tr>
<tr>
<td>Chlorine</td>
<td>color and odor (no equation)</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>slight explosion (no equation)</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>fumes in moist air (no equation)</td>
</tr>
<tr>
<td>Hydrogen sulfide</td>
<td>odor; turns lead acetate paper black</td>
</tr>
<tr>
<td>Sulfur dioxide</td>
<td>odor; decolorizes permanganate</td>
</tr>
<tr>
<td>Oxygen</td>
<td>splint burns (no equation)</td>
</tr>
<tr>
<td>Nitric oxide</td>
<td>oxidation to brown fumes</td>
</tr>
<tr>
<td>Nitrogen dioxide</td>
<td>color and odor (no equation)</td>
</tr>
<tr>
<td>Water</td>
<td>anhydrous copper sulfate → blue</td>
</tr>
<tr>
<td>Starch</td>
<td>iodine (no equation)</td>
</tr>
<tr>
<td>Hydroxyl ion</td>
<td>litmus (no equation)</td>
</tr>
<tr>
<td>Chloride</td>
<td>silver nitrate and nitric acid</td>
</tr>
<tr>
<td>Sulfate</td>
<td>barium chloride</td>
</tr>
<tr>
<td>Nitrate</td>
<td>brown ring</td>
</tr>
<tr>
<td>Phosphate</td>
<td>molybdate</td>
</tr>
<tr>
<td>Carbonate</td>
<td>acid and lime water</td>
</tr>
<tr>
<td>Sulfide</td>
<td>odor; lead acetate paper</td>
</tr>
<tr>
<td>Bromide</td>
<td>chlorine water + $\text{CCl}_4$</td>
</tr>
<tr>
<td>Iodide</td>
<td>chlorine water + $\text{CCl}_4$</td>
</tr>
<tr>
<td>Sodium ion</td>
<td>flame yellow (no equation)</td>
</tr>
<tr>
<td>Potassium ion</td>
<td>flame lavender (no equation)</td>
</tr>
<tr>
<td>Calcium ion</td>
<td>flame brick red (no equation)</td>
</tr>
<tr>
<td>Ion</td>
<td>Test</td>
</tr>
<tr>
<td>---------------------</td>
<td>-----------------------------------------------</td>
</tr>
<tr>
<td>Barium ion</td>
<td>green flame test (no equation)</td>
</tr>
<tr>
<td>Lithium ion</td>
<td>red flame test (no equation)</td>
</tr>
<tr>
<td>Zinc</td>
<td>white precipitate with $\text{H}_2\text{S}$</td>
</tr>
<tr>
<td>Arsenic</td>
<td>yellow precipitate with $\text{H}_2\text{S}$</td>
</tr>
<tr>
<td>Antimony</td>
<td>orange precipitate with $\text{H}_2\text{S}$</td>
</tr>
<tr>
<td>Copper</td>
<td>black precipitate with $\text{H}_2\text{S}$; blue complex with $\text{NH}_4\text{OH}$</td>
</tr>
<tr>
<td>Cadmium</td>
<td>yellow precipitate with $\text{H}_2\text{S}$</td>
</tr>
<tr>
<td>Ferric iron</td>
<td>blue with potassium ferrocyanide</td>
</tr>
<tr>
<td>Ferric ion</td>
<td>blue with potassium ferrocyanide; red with potassium thiocyanate</td>
</tr>
<tr>
<td>Silver</td>
<td>precipitate with chloride</td>
</tr>
<tr>
<td>Lead</td>
<td>precipitate with chloride; test with hot water</td>
</tr>
<tr>
<td>Mercourous ions</td>
<td>precipitate with chloride; test solubility</td>
</tr>
</tbody>
</table>
APPENDIX D

SEMIMICRO DEDUCTIVE LABORATORY MANUAL

by

SISTER ERNESTINE MARIE O'CONNELL

Published by

Washington Irving Publishing Company

Newark, New Jersey
TO THE TEACHER

The teacher should feel no hesitancy about adopting this manual. Although it is new, it involves the old. The familiar methods which have been taught and retaught throughout the years will be found here. The manipulations of semi-micro techniques are readily acquired, and once learned, they give the teacher the opportunity to devise many experiments other than those included here. No new laboratory is needed, and no new equipment is mandatory.

The order of the experiments in this manual is flexible and permits its use with any standard high school chemistry text. What one may choose, another may eliminate, but sufficient material is included to permit a good choice for all. In the back of the book there are also sets of drill problems and equations of the major types involved in high school work.

From the pedagogical aspect, the outline report forms have been arranged so as to call the students’ attention to a methodical presentation of results. It is to be hoped that the constant repetition involved in the use of this form will instill into the student a feeling for the scientific method of approach and report. It is a departure from the overused one-word answer system which many teachers feel is, in large measure, responsible for our high school “cook-book” chemists.

The author expresses the hope that this new book will meet with the approval of forward-looking high school teachers of the country in the field of chemistry. All inquiries, suggestions and criticisms are welcome and whole-hearted cooperation is extended to those who are about to embark on the course.
EXPERIMENT 1

THE BUNSEN BURNER OR MICRO BURNER

I: PRINCIPLE OF THE BURNER

(Note: These directions are for the gas burner. If there is no gas available, it is suggested that the instructor include a special experiment using that type of fuel and burner which is at hand.)

Note the construction of the burner, taking careful notice of the air vent and the small pin-hole opening for the gas. The opening for the air may be closed by revolving the burner stack or the perforated ring attached to the stack or by some similar device. Discover which your burner has.

Connect the burner to a source of gas. With the air-vent open, turn on and regulate the gas so as to obtain a small flame. The flame obtained should be blue and made up of two cones: the inner, greener cone, which is the reducing zone; and the outer, more colorless cone, which is the oxidizing zone. The tip of the inner cone is the hottest part of the flame. Try to light a burnt match stick in the lower cone.

Shut off the gas and allow the burner to cool, then close the air vent. Again turn on the gas and light it. Hold a piece of glass tubing in the upper part of this flame and observe the deposit that forms on it.

II: PRINCIPLE OF THE DAVY SAFETY LAMP

(This is an extra experiment for those who have time.)

Sir Humphry Davy, in 1815, studied the causes and prevention of firedamp explosions in mines. The explosions were caused by the ignition of methane and air. He found that if a flame is cooled, it is extinguished, and that before a gas will burn or explode, it must reach its ignition point.

Lower a closed spiral of thick copper wire over a candle flame and the flame will be extinguished. Copper is a good conductor of heat and lowers the ignition point of the combustible gases of the candle.

Place a fine wire gauze over a Bunsen flame. The flame at first does not pass through because of the heat conduction of the gauze. That gas is passing through may be determined by holding a lighted match above the gauze. If the gauze is held in the flame, the gas above the gauze will eventually burn. Repeat until the flame is obtained above the gauze; raise the gauze and observe that the flame goes out.
REPORT FORM FOR EXPERIMENT 1

The Bunsen Burner

I: Make a large sketch of the burner and label all its parts.

II: Make a large sketch of the flame when the air is admitted and label all its parts.

III: Sketch the luminous flame resulting when air is excluded from the burner.
I: DISCUSSION

In bending or drawing glass tubing, correct results are obtained only when the glass is rotated in the top part of the flame and is heated long enough to make the glass bend very easily and seem as if it will break or drop in two parts. Also, in bending tubing, the glass should be rotated, not twisted, during heating.

II: MANIPULATION

A: Bending glass tubing: Using a file or an ampoule cutter, make a scratch (do not saw) about 6 in. from the end of a piece of glass tubing. Pick up the tubing, placing the thumbs opposite and on both sides of the scratch, and curl the fingers over the tubing. Break the tubing as you would break a piece of wood (Fig. 1).

Fig. 1

After the tubing has been broken, adjust a wing tip on your burner and regulate the burner so as to obtain a luminous flame; then rotate the tubing in the flame (Fig. 2). During the rotation, heat the same area of the tubing until it becomes soft and readily workable. At this stage remove the tubing from the flame and bend it to an obtuse angle as in the diagram (Fig. 3). After the proper bend has been obtained, set the tube on an asbestos mat to cool. When cool, file the ends of the tubing until all rough pieces are chipped away, and then heat both ends in the flame to obtain a smooth, fire-polished finish. Repeat, using 10- and 12-in. pieces of tubing and make the figures illustrated in Figs. 3, 4, and 5.
B: *Making a medicine dropper or dropping pipet.* Cut a piece of glass tubing (6 mm outside diameter) 6 in. long. Without using a wing tip, heat the tubing in the center and rotate until the tubing is very soft. Remove from the flame and draw the tubing apart until a long, narrow capillary portion is obtained (Fig. 6). Place the tubing on an asbestos mat to cool. Be careful never to put hot tubing in direct contact with the desk top since the cold surface will cause the tubing to crack. After the tubing has been cooled, cut the capillary portion into two sections, 3 in. long. Now take one of these sections and heat the wide end until it becomes very soft; remove from the flame and press down on an asbestos mat. This last operation will give a flared end over which the rubber bulb may be fitted. Again allow to cool; fire-polish the small end by heating a very short time in the flame; attach the rubber bulb. Save the other end of the drawn tube for use in making the wash bottle.

C: *Making a spatula and crusher.* Cut a glass rod 6 in. in length. Heat one end until very soft. Remove from the flame and immediately press the heated end vertically on an asbestos mat or charcoal block. This gives a rounded flat end. Now, heat the other end and press it at a 45° angle against a mat or block. This gives the opposite end of the rod as shown in the diagram (Fig. 7).

![Diagram of a spatula and crusher](image)

D: *Making a wash bottle.* If a 250-ml Florence flask is used, make an obtuse angle bend in a piece of glass tubing 5 in. long. Then take a foot of tubing and make an acute angle supplementary to the obtuse angle already constructed. (A ruler should lie right along tubes A and B in the drawing.) Fit this latter bend through a two-hole rubber stopper which fits the flask in use and gauge the place where the lower bend must be made. Make a suitable bend here. Cut the dropping pipet, prepared above, to about an inch in length and attach it with a rubber tubing section to the acute angle B. Fire-polish all rough ends and assemble as in the diagram (Fig. 8).
EXPERIMENT 3

USE OF THE METRIC SYSTEM —
COMPARISON OF METRIC AND ENGLISH SYSTEMS

I: MEASUREMENT OF LENGTH

Using a centimeter ruler, measure the length, width, and thickness of a test tube block and a glass plate, recording the results in centimeters and inches on the report form one page beyond. Now, exchange a similar set of objects with a neighboring student and again measure and record your results. After these measurements have been completed, compare your results with those obtained by your neighbor.

II: MEASUREMENT OF VOLUME

Add water to a graduated cylinder to some point within the lines or graduations. Place the cylinder on a flat surface and make a reading of the volume in milliliters and in cubic centimeters. In making the reading, read the lowest part of the curved surface or meniscus at the top of the liquid and keep your eye on a level with that surface (Fig. 9).

![Fig. 9](image)

Now fill a test tube and pour water into the graduated cylinder and read and record this volume. Do likewise, using a vial and a dropping pipet. Exchange these pieces with your neighbor and compare results as you did before.

With your test tubes, practice filling them to amounts of 1, 2, 3, 4, and 5 ml. Learn to estimate these quantities by remembering how far the tube is filled in each case. This will help speed your experiments later on in the course.

III: MEASUREMENT OF WEIGHT

Carefully mount the horn-pan balance on a clamp stand and see that the pointer is at zero. If it is not, notify your instructor.

Before making any weighings, observe these points:

1. Be sure the pointer is at zero.
1. Be sure the pointer is at zero.
2. Have both pans free from dust and other particles.
3. When adjusting weights, always use a forceps. DO NOT PICK UP WEIGHTS WITH FINGERS.
4. Place the object to be weighed in the left hand pan and the weights in the right.
5. Never place corrosive substances directly in the scale pan. Place paper or a glass plate in the pans first and adjust to zero.
6. After completing a weighing, place the weights back in the box.
7. Record all weights in your laboratory notebook—never on scraps of paper.

After having read the foregoing directions carefully, weigh the following pieces of apparatus:

   a: crucible
   b: test tube
   c: crucible + test tube

Now weigh out 0.2 g of salt in the following manner: Place a small piece of paper in the left pan of the balance and add weights to restore the pointer to zero. If desired, a similar piece of paper may be placed in the right pan as counterbalance, thus avoiding the use of extra weights. Now place a 0.2 g weight in the right pan and add salt to the paper in the left pan until both pans almost balance. Perhaps you can get exactly the right amount of salt. If not, add weights to the right pan until perfect balance is obtained.
REPORT FORM FOR EXPERIMENT 3

Metric System

<table>
<thead>
<tr>
<th>Object in set 1</th>
<th>Length cm</th>
<th>Length in.</th>
<th>Width cm</th>
<th>Width in.</th>
<th>Thickness cm</th>
<th>Thickness in.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

in set 2

<table>
<thead>
<tr>
<th>Apparatus</th>
<th>Volume ml</th>
<th>Volume liters</th>
<th>Apparatus</th>
<th>Volume ml</th>
<th>Volume liters</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Object | Weight in mg | Weight in g | Weight in kg
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

There are ................ cm in 1 inch. THIS NUMBER SHOULD BE MEMORIZED.
SUGGESTIONS TO STUDENTS

Perform all experiments while holding equipment at arm's length. Equipment is small and simple, no vessel being over 25-ml capacity. When a test is to be made on a liquid, the solution to be tested is placed in a 5- or 10-ml test tube, filling it up to a depth of about 1 cm. To this is added, dropwise, the reagent.

To determine odor, waft the fumes toward the nose with the cupped hand; NEVER INHALE DIRECTLY FROM A TUBE OR BOTTLE (Fig. 10).

Transfer of solids: Never pour directly from a bottle to a flask or test tube, but transfer the solid to a small piece of creased, glazed paper, and, using this as a slide, slip the solid into the test tube (Fig. 11).

Transfer of liquids: Turn the palm of the hand upward and grasp the stopper of the reagent bottle with the index finger and third finger and remove it. Holding the stopper between the aforementioned fingers, grasp the bottle and pour the liquid into the tube. Small amounts of liquids may be transferred by means of a dropper or pipet (Fig. 12).

Measuring volume: Graduated cylinders, burettes, and pipets are used to measure the volume of liquids. In reading the level of the liquids in these tubes, keep the eye on a level with the bottom of the meniscus and read the line upon which it rests.

Testing with indicator paper: Transfer a drop of the solution to be tested by means of a stirring rod to the test paper.

Heating materials: This is always done in Pyrex, thin-walled glass--
ware or porcelain. Thick-walled bottles are not heat resistant. When heating test tubes, hold them at an angle pointing away from you.

Filtration: Support a funnel in a ring stand and touch the stem to the wall of the receiving vessel. Fold the filter paper in halves along the diameter, and again in halves. Open the quarter nearest you and place the cone in a moist funnel, easing it in until it fits snugly (Fig. 13).

Centrifugation: The material to be centrifuged must be placed in a tube which fits the centrifuge and must be counterbalanced by a similar tube containing water. A half minute generally suffices to settle any precipitate.

Solubility of gases: The tube of gas is inverted in a beaker of water. After a few minutes, water will appear in the tube if the gas is dissolving.

Solubility of solids: Any evidence of dissolving action, like the evolution of a gas or the disappearance of the solid, is taken as indicative of solubility.

Combustion testing: A burning splint inserted in a gas will either burn or go out, and the gas may ignite or not. If the gas burns, efforts should be made to identify the products of combustion.

Density: This may be inferred from the type of apparatus recommended in the collection of the gas or may be calculated. Relative density is compared with that of air taken as 1.29.

PRACTICE OF COMMON LABORATORY OPERATIONS

I: To a 4-in. test tube, add about a gram (a spatulaful) of a ferric salt; add 5 ml K<sub>2</sub>CrO<sub>4</sub> solution and 2 drops of conc. NH<sub>4</sub>OH. Keep the tube corked.

(Note: When directions call for a “spatulaful” they mean the amount of solid that can be easily held on the tip of a microspatula.)

II: EVAPORATION: Place a few drops of the mixture prepared in I in a crucible or evaporating dish and heat GENTLY until the liquid almost disappears. Overheating will cause spattering of the solid.

III: CENTRIFUGATION: Stir the original mixture and quickly pipet off about 10 drops into a centrifuge tube half filled with water. Fill another tube to the same extent with water and place both tubes in the centrifuge. Let the machine run for about half a minute, then remove both tubes.
IV: PRECIPITATION: To 2 ml of H₂O in a centrifuge tube, add about 8 drops of the centrifugate obtained in III and add 2 ml of Pb(NO₃)₂. Balance this tube with one of water and centrifuge.

V: DISTILLATION: Pour half the remainder of the original mixture into a 4-in. test tube fitted with a one-hole stopper and delivery tube leading to a small test tube. Place this second tube in a beaker of ice. Heat the mixture until 1 ml of distillate is collected. Test it with litmus paper. (CAUTION: If a Bunsen burner is used, the heat must be applied very gently by holding the burner in the hand and moving the flame over the tube. If a micro burner is used, less caution is needed.)

VI: FILTRATION: Fold a filter paper in quarters and slip into a funnel. Stir the remainder of the original mixture and quickly pour it or pipet it into the filter. Catch the filtrate in a small test tube or on a watch glass.

VII: The following three preparations are for gases. Facility in gas collection will be a great asset later in the course.

Gas Set-up A (Fig. 15) for gases insoluble in water.

Place a small piece of mossy zinc in a 4-in. test tube fitted with a stopper and conventional delivery tube (Fig. 5). Fill the pneumatic trough with water and immerse 5 4-in. test tubes or collection vials in the trough. Add about 3 ml of dil. H₂SO₄ to the zinc and insert the delivery tube so that the end is beneath the water in the trough. Slip a collecting vial over the end of the delivery tube and allow it to fill with gas. When the vial is full of gas (empty of H₂O) remove it and replace it with another vial. In removing full vials, cover the end of the vial with the thumb or insert the cork while holding the tube mouth under water.
Gas Set-up B (Fig. 16) for gases lighter than air and soluble in H₂O.

To a generator tube (a 4-in. test tube) add about a gram of NH₄Cl (a spatula-ful) and an equal amount of Ca(OH)₂. Attach an oblique angle delivery tube and heat the generator gently. Collect gas in vials by the upward displacement of air. A piece of moist litmus paper placed at the end of the collecting vial will turn color when the vial is filled.

Gas Set-up C (Fig. 17) for gases heavier than air and soluble in H₂O.

To about a gram of salt, NaCl, in a generator tube (a 4-in. test tube) add 3 ml of dil. H₂SO₄ and attach the acute angle delivery tube made in Glass Manipulation (Fig. 4). Collect the gas by the downward displacement of air.
GENERAL OUTLINE REPORT FORM

A: OCCURRENCE
1: Place and state
2: Method of extraction (if any)

B: PREPARATION
1: Laboratory (description + equations) and diagram
2: Commercial (description + equations)

C: PROPERTIES
1: Physical
   a: density
   b: color
   c: odor
   d: taste
   e: solubility
   f: change of state (if any)
   g: special (allotropes, isotopes, etc.)
2: Chemical
   a: stability
   b: combustibility
   c: reactions with elements
   d: reactions with compounds
   e: special (oxidizing or reducing agent, acid, base, etc.)
   f: test for positive identification

D: STRUCTURE
1: atomic diagram
2: molecular diagram
3: ionic diagram

   a: placement in periodic table
   b: metal or non metal; electron donor or acceptor

E: USES

F: OTHER IMPORTANT EQUATIONS
EXPERIMENT 4

PHYSICAL AND CHEMICAL PROPERTIES OF MATTER

Some of the physical properties of matter used to identify certain substances are: color, odor, taste, and solubility in water or other solvents. Some of the chemical properties similarly used are: combustibility, ease of corrosion (for metals), reactions with acids and other compounds, reactions with elements.

These properties fall under heading "C" in your General Outline Report Form and have subheadings C-1 Physical Properties and C-2 Chemical Properties. Hence, maintaining the lettering system that is used in the outlines, we shall proceed:

C-1 PHYSICAL PROPERTIES

C-1a Density Calculate the density of iron, hydrogen and sulfur.
C-1b Color Examine the color of sulfur, iron, salt and water.
C-1c Odor Determine the odor of NH₄OH and of HCl and of sulfur.
C-1d Taste Note: DO NOT SWALLOW THE TASTED SUBSTANCE. NEVER TASTE SOLUTIONS UNLESS DIRECTED TO DO SO.

Place a crystal of salt in a test tube and add a few drops of H₂O. With a clean glass rod, help the crystals to dissolve and then touch the tip of the rod to your tongue. Fill a small test tube half full of H₂O and add a drop of HCl. Stir with a glass rod and taste as before.

C-1e Solubility Place a crystal of sugar on a glass plate and add a drop of H₂O. Repeat, using a crystal of K₂CrO₄. Let stand a few minutes. If the crystal remains, remove it with a pair of forceps. Note the color of the water. Evaporate the water and examine the residue.

C-2 CHEMICAL PROPERTIES

C-2a Stability Note the odor of the gas escaping from a bottle of HCl and from NH₄OH. This is evidence that the compound is decomposing and that the solution is not too stable.
C-2b *Combustibility* Place a spatulaful of red phosphorus on an asbestos mat and ignite it. Place a small piece of zinc in a test tube and add 5 drops of dil. HCl. Test the escaping gas with a lighted splint.

C-2c *Reactions with elements* Place a spatulaful of granulated zinc in a crucible and heat gently for about 5 min. The white powder formed is ZnO.

Hold a piece of magnesium ribbon in the flame by means of a forceps and ignite it. The white powder formed is MgO.

Place some powdered sulfur on a penny and hold the coin in the flame of a burner. Remove and note the black color of the coin face. This is CuS.

C-2d *Reactions with compounds* Half-fill a test tube with H$_2$O and add 2 drops of HCl. Drop into this a small piece of zinc.

Repeat, using a piece of tin instead of the zinc, and if time permits, try some other metals.

C-2e *Special properties* In a small test tube place 3 ml of H$_2$O and add 2 drops HCl. Test the solution with litmus paper and with phenolphthalein.

In another test tube, place 2 ml of H$_2$O and add 2 drops of NaOH solution. Test this also with litmus paper and then add a drop of phenolphthalein.

In a small test tube containing 3 ml H$_2$O, add 1 drop KMnO$_4$. Drop in a crystal of Na$_2$SO$_3$ and shake well. This action is caused by electrons transferring from one substance to another and is called oxidation-reduction. There are many "redox" reactions that you will learn later. The power to oxidize or to reduce is one of the special properties that we test for.
EXPERIMENT 5
ELEMENTS, MIXTURES and COMPOUNDS

I: ELEMENTS

Examine a spatulaful of iron powder and one of flowers of sulfur, to determine their physical properties. [C-1a, C-1b, C-1c, C-1d, C-1e and C-1g (magnetic effect).]

Using your burner and a pinch of the element on a metal spatula, determine the combustibility of each [C-2b].

II: MIXTURES

**EXTINGUISH FLAMES BEFORE WORKING WITH CS₂**

Place a spatulaful of powdered sulfur and one of powdered iron on a piece of paper. Intimately mix them with a stirring rod and examine the mixture with a hand lens. Bring a magnet to the UNDER SIDE of the paper containing the mixture and observe the effect [C-1g]. Now, transfer about one-third of the mixture to a centrifuge tube and add 2 ml CS₂. Shake thoroughly, centrifuge, and draw off the supernatant liquid onto a watch glass. Set it aside to evaporate [C-1e].

III: COMPOUNDS

Again place a spatulaful of sulfur and a little less of powdered iron in a crucible. Cover and heat the mixture. After about 3 minutes, examine the residue with a lens.

After the residue has cooled, apply a magnet to it. Add a drop of CS₂ and observe the result. After this has evaporated, add a drop of HCl and observe the residue and the odor of the gas evolved [C-2d].

Add a drop of HCl to separate portions of S and Fe [C-2d].

**OPTIONAL EXPERIMENT FOR THOSE WHO HAVE TIME**

**THE SEPARATION OF CREAM FROM MILK:**

To a centrifuge tube, add 1 ml milk. Next add gently 1 ml conc. H₂SO₄ and stir the contents with a stirring rod and mix thoroughly. Centrifuge; add 10 drops of very hot water and centrifuge again. The fat or cream forms a separate layer on top. This method is used commercially to determine the percentage of butterfat in milk or cream, and is known as the Babcock Method.
REPORT FORM FOR EXPERIMENT 5
Elements - Mixtures - Compounds

C-1 Physical Properties:

<table>
<thead>
<tr>
<th></th>
<th>A (IRON)</th>
<th>B (SULFUR)</th>
<th>C (IRON + SULFUR)</th>
<th>D (IRON SULFIDE)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-1a</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-1b</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-1c</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-1d</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-1e</td>
<td></td>
<td></td>
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<td></td>
</tr>
<tr>
<td>C-1g</td>
<td></td>
<td></td>
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</tbody>
</table>

C-2 Chemical Properties:

<table>
<thead>
<tr>
<th></th>
<th>A (IRON)</th>
<th>B (SULFUR)</th>
<th>C (IRON + SULFUR)</th>
<th>D (IRON SULFIDE)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-2a</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-2b</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C-2c</td>
<td></td>
<td></td>
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<tr>
<td></td>
<td></td>
<td></td>
<td>Fe + S →</td>
<td>FeS + HCl →</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>FeS + CS₂ →</td>
</tr>
<tr>
<td>C-2d</td>
<td></td>
<td>Fe + HCl →</td>
<td>S + HCl →</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>S + CS₂ →</td>
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</tbody>
</table>
EXPERIMENT 8

CRYSTALS AND CRYSTALLIZATION

I: WATER OF CRYSTALLIZATION (or water of hydration)

Heat small crystals of the following in separate test tubes. If any fumes appear to escape, test them with cobalt nitrate paper to prove whether or not they contain water.

(a) copper sulfate  (c) sodium chloride
(b) sodium carbonate  (d) potassium chlorate

Exposure crystals of CaCl₂ and Na₂CO₃ on a glass plate until the end of the laboratory period. If no marked change occurs, label them and put them in your desk until the next laboratory session.

II: FORMATION OF CRYSTALS (suggested demonstration)

Almost fill a small test tube with a hot, concentrated, saturated solution of K₂Cr₂O₇ or CuSO₄. Place a string in the test tube and set it aside to cool.

III: CRYSTALLINE ALLOTROPES

Dissolve a tiny bit of sulfur in a few drops of CS₂. Pour the solution onto a glass plate and let the solvent evaporate. Examine the shape of the resulting crystals.

In an old test tube, place some flowers of sulfur and heat the tube gently. Once the sulfur is all molten, pour it quickly into a small cone of filter paper. As soon as the top begins to harden, open out the filter paper and observe the shape of the crystals formed in this way.

Again place flowers of sulfur in the test tube and again heat the tube slowly and gently, noticing the change in color and consistency which the sulfur undergoes. These colors and thickness changes indicate the presence of different liquid allotropic forms. When the liquid sulfur boils, pour it into a dish or beaker of cold water and observe the form of sulfur resulting from this treatment.
REPORT FORM FOR EXPERIMENT 6
Crystals and Crystallization

I: Water of crystallization: Write equations for all reactions that occurred, and indicate all color and shape variations.

a:

b:

c:

d:

e:

f:

II: Crystalline allotropes: Tabulate the results of the experiment in the form below.

<table>
<thead>
<tr>
<th>Crystalline Form</th>
<th>Temperature</th>
<th>Color</th>
<th>How obtained</th>
</tr>
</thead>
<tbody>
<tr>
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</tbody>
</table>
EXPERIMENT 7

PREPARATION AND PROPERTIES OF OXYGEN

What weight of O₂ will be formed by the complete decomposition of 10 g of KClO₃? What percent of KClO₃ is O₂? What percent of HgO is O₂? Which is the better source on this basis?

I: LABORATORY PROCEDURE

Mix about 1 g KClO₃ with half as much MnO₂ on a piece of paper. Transfer the mixture to a 4-in. test tube fitted with a number 00 one-hole stopper and a type A delivery tube (see gas set-up A, p. 10) leading to a trough of water in which 5 collecting vials are submerged.

Holding the burner, carefully heat the test tube, keeping the burner in constant motion. (CAUTION: NEVER HOLD THE HAND DIRECTLY UNDER A HEATED OBJECT AT ANY TIME.) Let the first few bubbles escape to rid the generator and tubes of air. Now, place the thumb over the mouth of one of the vials filled with water and bring it directly over the end of the delivery tube, still under water. The bubbles of O₂ will enter the vial and displace the H₂O. Control the heat so that a steady flow of O₂ results. If the gas is being generated too quickly, remove the heat; however, if the water begins to suck back up the delivery tube, either withdraw the tube or apply more heat. As soon as all the water is displaced from a vial, as shown by the bubbles of gas in the trough, quickly stopper it under water and replace it with another vial of water. Collect the O₂ until all 5 vials are filled.

II: PROPERTIES

Calculate the density of the O₂. [C-1a]

Note the color, odor and taste. [C-1b, C-1c, C-1d] Invert a vial of O₂ in a beaker of water. [C-1e]

Insert a glowing splint into a vial of gas and observe the results. [C-2b, C-2f]

Heat some powdered charcoal on a spatula until it glows and quickly insert this into a vial of O₂. [C-2c]

Repeat the above, using powdered sulfur instead of charcoal.

III: REACTION

Test the material remaining in the generator by dissolving a bit of it in H₂O. Filter or centrifuge the resulting suspension and to the clear liquid add a drop of AgNO₃. The formation of a curdy, white precipitate is a positive test for the Cl⁻ ion. [B-1 proof]
A: OCCURRENCE

1: Place and state
   a: 
   b: 

2: Methods of extraction
   a: 
   b: 

B: PREPARATION

1: Laboratory
   a: from a chlorate
   b: from sodium peroxide.

2: Commercial
   a: from air
   b: from water.

C: PROPERTIES

1: Physical
   a: 
   b: 
   c: 
   d: 
   e: 
   f: 
   g: allotropic forms.

2: Chemical
   a: 
   b: 
   c: \( S + O_2 \rightarrow \)
   C + O_2 \( \rightarrow \)
D: STRUCTURE

1: atomic diagram  2: molecular diagram  3: ionic diagram

a: Placement in Periodic Table?
b: Metal or non-metal?  Electron donor or acceptor?

E: USES

F: OTHER IMPORTANT EQUATIONS

\[
\begin{align*}
H_2 & + O_2 \rightarrow \\
P & + O_2 \rightarrow \\
Mg & + O_2 \rightarrow \\
N_2 & + O_2 \rightarrow \\
NO & + O_2 \rightarrow \\
HgO & \rightarrow \\
H_2O & \rightarrow \\
ZnS & + O_2 \rightarrow \\
Zn & + O_2 \rightarrow \\
Cu & + O_2 \rightarrow \\
Fe & + O_2 \rightarrow \\
H_2S & + (limited) O_2 \rightarrow \\
H_2S & + (excess) O_2 \rightarrow \\
CH_4 & + O_2 \rightarrow \\
C_2H_2 & + O_2 \rightarrow \\
CO & + O_2 \rightarrow \\
\end{align*}
\]
EXPERIMENT 8
PREPARATION AND PROPERTIES OF HYDROGEN

What weight of hydrogen will result from a complete reaction of 5 g of zinc? What volume will this occupy? If the gas is generated at 350°C and 756 mm pressure, what volume will the gas occupy at STP?

I: LABORATORY PROCEDURE

Prepare a gas generator, type "A" (Fig. 15) and have 5 collecting vials submerged in the collecting trough. Make sure you have stoppers at hand to fit these collecting tubes. Insert a bit of mossy zinc into the generator tube and add 2 ml dil. HCl. Insert the delivery tube and collect the gas by the displacement of H₂O. As each vial of gas is filled, stopper it immediately under water and stand it upside down in your test tube rack or on the table. [C-1a]

II: PROPERTIES

Physical

Observe the color, odor and taste of the gas. [C-1b, C-1c, C-1d]
Calculate the density of the gas. [C-1a]
Invert a tube of H₂ in a beaker of water. [C-1e]

Chemical

Bring a lighted splint to the mouth of the delivery tube and ignite the gas as it escapes. [C-2b]

Bring a lighted splint to the mouth of one vial of H₂. [C-2f]

Light the burner. Invert a vial of H₂ and bring it mouth to mouth with a vial of air in so doing. Now, bring each vial in turn to the burner flame, mouth first. [C-2c]

Invert a vial of H₂. Take out the cork and immediately insert a lighted splint. See if you can withdraw the splint and relight it as it passes through the burning H₂ at the mouth of the tube.

III: REACTION

To prove the reaction occurred, you may test for the chloride ion in the solution of the residue of the generator. The positive test for zinc ions may also be used and this consists in passing H₂S gas through a solution of the ions, and, if present, zinc will produce an insoluble white sulfide.
OUTLINE REPORT FORM FOR EXPERIMENT 8
Preparation and Properties of Hydrogen

A: OCCURRENCE
1: Place and state
2: Methods of extraction

B: PREPARATION
1: Laboratory
   a: metal + acid
   b: active metal + water
2: Commercial
   a: from water
   b:
   c:

C: PROPERTIES
1: Physical
   a:
   b:
   c:
   d:
   e:
   g: Isotopes
2: Chemical
   a:
   b:
   c: \( \text{H}_2 + \text{O}_2 \rightarrow \)
   e: \( \text{H}_2 \) is a good reducing agent.
D: STRUCTURE
1: atomic diagram
2: molecular diagram
3: ionic diagram

a: Placement in Periodic Table?
b: Electron donor or receiver? Metal or non-metal?

E: USES

F: OTHER IMPORTANT EQUATIONS:
\[ \text{N}_2 + \text{H}_2 \rightarrow \]
\[ \text{CuO} + \text{H}_2 \rightarrow \]
\[ \text{CO} + \text{H}_2 \rightarrow \]
\[ \text{Fe} + \text{HCl} \rightarrow \]
\[ \text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \]
\[ \text{Na} + \text{H}_2\text{O} \rightarrow \]
\[ \text{K} + \text{H}_2\text{O} \rightarrow \]
\[ \text{Ca} + \text{H}_2\text{O} \rightarrow \]

G: DIAGRAM OF LABORATORY PREPARATION.
EXPERIMENT 9
PHYSICAL AND CHEMICAL PROPERTIES OF WATER

I: PHYSICAL PROPERTIES

Taste a little distilled water and also note its color and odor. [C-1b, C-1c, C-1d] Place 1 ml of H₂O in each of three test tubes. To the first, add 5 drops of ether; to the second, 5 drops of alcohol; and to the third, 5 drops of carbon tetrachloride. This will give you some idea of the varying solubilities of liquids in water. [C-1e]

Place separate spatulafuls of sugar and PbCl₂ in separate test tubes. Add 5 drops of H₂O to each. After observing the solubility of the substances in the cold, heat the tubes gently. Repeat, using alcohol or CCl₄ as solvent.

II: CHEMICAL PROPERTIES

Into 2 ml of H₂O in a test tube, add a piece of calcium the size of a pin head. [C-2c] Add a drop of phenolphthalein to the resulting liquid.

Repeat the above, using iron instead of calcium. [C-2c]

Add a pinch of CaO (unslaked lime) to 2 ml of water in a test tube. [C-2d] Add a drop of phenolphthalein to this solution.
OUTLINE REPORT FORM FOR EXPERIMENT 9

Water

A: OCCURRENCE

C: PROPERTIES

1: Physical
   a:
   b:
   c:
   d:
   e: in ether
      in alcohol
      in CCl₄
   f:

2: Chemical
   a:
   b:
   c: Ca + H₂O →
      Fe + H₂O →
   d: CaO + H₂O →
D: STRUCTURE
   1: Molecular diagram

E: USES

F: OTHER IMPORTANT EQUATIONS
   Water purification: $\text{Cl}_2 + \text{H}_2\text{O} \rightarrow$
   Hardening of mortar: $\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow$
   Softening of hard water: $\text{Ca(HCO}_3\text{)}_2 \rightarrow$
   $\text{H}_2 + \text{O}_2 \rightarrow$
   $\text{HCl} + \text{NaOH} \rightarrow$
   $\text{H}_2\text{SO}_4 + \text{KOH} \rightarrow$
EXPERIMENT 10
METHODS OF PURIFYING WATER

I: DISCUSSION

Four methods of purifying water are: boiling, distillation, aeration, and filtration. Boiling and distillation kill bacteria, but distillation also separates organic and inorganic matter. Both of these methods are too expensive, however, to be employed on a large scale. Aeration has the drawback of oxidizing the food of the bacteria and leaving them in the water to die. The most common method in use for the purification of water in large cities is filtration. The water is allowed to filter through sand and gravel beds which hold back the solid material. If an aluminum salt is added to the water, it hydrolyzes, forming Al(OH)_3, a gelatinous precipitate. This precipitate carries with it not only solid material, but also bacteria. The addition of chlorine water in small amounts generally destroys all remaining bacteria.

II: EXPERIMENT

To 2 ml of muddy H_2O in a test tube, add a crystal of alum (aluminum potassium sulfate) and shake until the salt dissolves. Let the tube stand until the contents settle. This is purification by coagulation.

Place about 3 ml of muddy H_2O in a 4-in. test tube fitted with a type "C" delivery tube leading to a vial immersed in ice. Boil the H_2O in the generator tube very gently until a few drops of H_2O vaporize and condense in the collecting vial. This is distillation.

If time permits, repeat, using 3 ml of tap water with a drop of NaOH added to it. Collect the distillate in three separate test tubes, a few drops to each. To each sample of distillate, add a drop of phenolphthalein to determine the amount of NH_3 being carried over into each consecutive portion.
REPORT FORM FOR EXPERIMENT 10

Water Purification

<table>
<thead>
<tr>
<th>PROCESS</th>
<th>REMOVES</th>
</tr>
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<tbody>
<tr>
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</tbody>
</table>
EXPERIMENT 11

ELECTROLYSIS OF WATER

Cut two 18-mm lengths of glass tubing, and seal one end of each piece. Force each through a one-hole rubber stopper or wrap it in asbestos paper until it has a large enough diameter to be easily grasped by a burette clamp. Now, add tap water to an evaporating dish until the dish is three quarters full, and to this water add a gram of Na₂SO₄. After the complete solution of the sulfate is obtained, fill the two tubes with this solution, by means of a fine pipet. Set up the apparatus as shown in the diagram, using electric or adhesive tape to insulate the connections between the copper wire and the platinum wire.

Allow the electrolysis to proceed for half an hour or until one of the tubes is about half filled with gas. Using a ruler, measure the height of the gas in each tube. Remove each tube and invert by placing the thumb over the opening. Shake the H₂O down into the bottom of the tube. Remove the thumb and test the gas with a glowing splint about the size of a toothpick.
EXPERIMENT 12
ELECTROMOTIVE SERIES

EXPERIMENT
Half fill six small test tubes with H₂O. Add 2 drops of conc. HCl to each. To the first, add a piece of zinc; to the second, a piece of copper shot; to the third, an iron nail; to the fourth, a piece of lead; and to the fifth, a strip of magnesium ribbon. In the last, place a piece of tin. Note the reaction in each case. Identify any products that you can.

Set up five test tubes in your rack. To the first, add a solution of AgNO₃ to a depth of about 1 cm. To the second, add a like amount of MgSO₄; to the third, ZnSO₄; to the fourth, CuSO₄; and to the fifth, NaCl. Put a piece of clean magnesium ribbon in each, and carefully note the effect. Repeat, using strips of clean zinc instead of magnesium. If time permits, repeat with fresh solutions and pieces of copper shot.

REPORT
Using the following table, report all physical and chemical changes observed and write the reactions that occurred.

<table>
<thead>
<tr>
<th>METAL</th>
<th>REACTION WITH HCl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zinc</td>
<td></td>
</tr>
<tr>
<td>Iron</td>
<td></td>
</tr>
<tr>
<td>Lead</td>
<td></td>
</tr>
<tr>
<td>Copper</td>
<td></td>
</tr>
<tr>
<td>Magnesium</td>
<td></td>
</tr>
<tr>
<td>Tin</td>
<td></td>
</tr>
</tbody>
</table>
Write the reactions for all chemical changes that occurred. Also, indicate all physical changes that were observable.

<table>
<thead>
<tr>
<th></th>
<th>MAGNESIUM</th>
<th>ZINC</th>
<th>COPPER</th>
</tr>
</thead>
<tbody>
<tr>
<td>AgNO₃</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgSO₄</td>
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<td></td>
<td></td>
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<tr>
<td>ZnSO₄</td>
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<td></td>
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<tr>
<td>CuSO₄</td>
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<td></td>
<td></td>
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<tr>
<td>NaCl</td>
<td></td>
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</tbody>
</table>
EXPERIMENT 13
SOLUTIONS

I: TYPES OF SOLUTIONS

A: *Gases dissolved in liquids.* Place 3 drops of dil. NH₄OH in a test tube. Warm it gently and waft the fumes toward your face. Also, hold in the fumes a piece of moist blue and red litmus paper.

Place 10 drops of H₂O in a test tube and pass some H₂S into the water. Shut off the gas supply and remove the test tube. Carefully smell the solution. By means of a pipet, carefully transfer a few drops of this solution to a test tube containing 1 ml Pb(NO₃)₂ solution.

B: *Liquids dissolved in liquids.* To 3 drops of H₂O in a test tube, add 1 drop of alcohol; shake; add several more drops of alcohol.

Place 3 drops of H₂O in a test tube and add 3 drops of kerosene. After adding the kerosene, shake the tube thoroughly and allow the mixture to stand until clear.

C: *Solids dissolved in liquids.* Add a few crystals of salt to 2 ml H₂O in a test tube. Observe the effect.

Place two separate spatulafuls of CuSO₄ in separate test tubes. Add 2 ml H₂O to each. Place one of the tubes in a reagent block and continually stir the other with a stirring rod. From time to time, compare the tubes. *(Note: Of course, if the lumps are large, the rate of solution is slower than if the substance is powdered.)*

Place 3 g of KN₀₃ in 2 ml H₂O in a test tube. Now, heat the tube gently until complete solution is obtained. Set the tubes aside to cool and observe any changes that appear.

II: REPORT

Write up this experiment in good English, using complete sentences and clearly describing the effects noticed. Compare the different agents as to their effect on solution rate.
EXPERIMENT 14
COLLOIDS AND EMULSIONS

I: COLLOIDS
A: PREPARATION OF COLLOIDS

1: Sols. Sols are liquid colloids which may be prepared in several different ways. Among them are:

a: Grinding: By means of a crusher, pulverize some clay, about the size of a pea, in a small test tube. Add about 2 ml H₂O and centrifuge the mixture. Observe that the centrifugate is not clear, yet no particles settle out and no individual particles of clay are visible.

b: Chemical action: Heat 1 ml of H₂O to boiling. Add 1 drop of FeCl₃ solution; the FeCl₃ will hydrolyze, producing colloidal Fe(OH)₃. Sols are usually recognized by their cloudy appearance. However, some sols are as clear as solutions, to the naked eye. (Save this sol for use later in the experiment.)

2: Gels. Gels are solid colloids among the commonest of which is gelatin. To 1 ml of cold H₂O in a test tube, add 3 spatulafuls of powdered gelatin. Let this stand for a minute and then heat until the gelatin dissolves. Cool under a running tap until the gel sets. Save this also for later use.

B: PRECIPITATION OF COLLOIDS

Many sols are charged negatively and can therefore be precipitated by neutralizing these charges whose mutual repulsion is responsible for the continued separation of the particles. A suitable electrolyte will perform this neutralization of charge and permit the particles to clump together and settle out. This happens when salt water meets muddy river water; a delta results.

To the Fe(OH)₃ sol prepared above, add a drop or two of Na₂SO₄ and shake the tube thoroughly.

To the gel obtained above, add a pipetful of concentrated salt solution.

II: EMULSIONS
A: Soap: Take two small test tubes and fill each to a depth of 1 cm with kerosene. Add an equal volume of H₂O and to one of the tubes add 5 drops of soap solution. Shake both thoroughly and allow them to stand a while before recording the results.
III: OPTIONAL PREPARATIONS

To 2 ml of Na₂SiO₃ in a test tube (diluted 10-1) add 10 drops of conc. HCl and stir.

Place 5 ml alcohol in a crucible and add 1 ml saturated solution of Ca(C₂H₃O₂)₂ and mix well. Let the mixture stand a minute and then remove it from the crucible to an asbestos mat and ignite the gel.

To half a spoon of gelatin, add 20 ml H₂O. Using two flexible wires, attach two copper rods to a 6-volt DC current. Tape the upper ends of the copper rods so that you can hold them, and rub the lower ends together under the surface of the gelatin-water solution. Some of the copper dislodged is of colloidal dimensions.

Use this space to report this experiment
EXPERIMENT 15
PERCENTAGE OF WATER IN COPPER SULFATE

I: PROCEDURE
The weight of the element or group of elements under consideration, compared with the weight of the whole compound, will give the ratio of the weights of these substances. This ratio, expressed in percent, gives the percent of the element desired. Hence, if we know that in 50 g of copper sulfate there are 20 g of copper, then we conclude the ratio of Cu to CuSO₄ is 20:50, or 2:5, which would be 40 percent. Experimentally we are going to determine the percent of water in crystals of hydrated copper sulfate. We shall weigh the initial crystals which contain water of hydration; we shall heat the crystals to drive off the water and then reweigh the dry crystals. Obviously, the weight loss is due to the water which was there and escaped. Knowing this figure, we can then calculate the percent of water in the original crystals.

II: DATA
Weight of crucible + hydrated CuSO₄ crystals
Weight of empty crucible
Weight of hydrated CuSO₄
Weight of crucible after heating
Weight of water lost
Percentage of water of hydration
Theoretical percentage of water of hydration
Percent error

DO CALCULATIONS HERE
EXPERIMENT 16
DETERMINATION OF THE MOLECULAR VOLUME OF OXYGEN
and
THE PERCENTAGE OF OXYGEN IN KClO₃

I. EXPERIMENT

Weigh a clean, dry, 4-in. test tube and add to it about 0.20 g of KClO₃. Add about 0.1 g MnO₂ and reweigh the test tube with its contents. Set up the apparatus as in gas set-up type A, except that, instead of a collecting vial, a 4-oz. bottle is substituted. Generate the gas until no more oxygen bubbles over; then, with the flame still under the generator, remove the delivery tube from the water. Remove the flame and let the generator cool to room temperature before reweighing it. Meanwhile, equalize the levels of the water inside the bottle and that outside in the trough. By the aid of a glass plate, remove the bottle of O₂ + H₂O and place it upright on the desk. Mark the level of the H₂O with a marking pencil. Fill the bottle to the brim with water, carefully recording the amount needed. This value is the volume of O₂ generated. Record the temperature of the water in the trough and the barometric pressure of the room.

II. DATA

Weight of tube + KClO₃
Weight of tube alone
Weight of KClO₃
Weight of tube + KClO₃ + MnO₂
Weight of tube + residue after generating oxygen
Weight of O₂ evolved
Volume of O₂ evolved
Temperature of the water
Barometric pressure
Volume of O₂ at STP
Volume occupied by 1 g of O₂
Molecular volume of O₂
Percentage of O₂ in KClO₃
Theoretical percentage of O₂ in KClO₃
Percent error
EXPERIMENT 17
THE DETERMINATION OF THE VALENCE OF MAGNESIUM

I: EXPERIMENT

Variations in the method of collecting the releasing hydrogen may be used. If eudiometer tubes are available, accordion-pleat a small piece of magnesium ribbon and insert it under an inverted funnel below an eudiometer tube which contains HCl, diluted 1:3 with water.

A 100-ml graduated cylinder may be used in place of the eudiometer, and the readings of the released hydrogen may be made directly on it.

Or a 4-oz. bottle may be inverted in a dish or pneumatic trough. In this case, weight out a piece of magnesium ribbon that is somewhat less than 0.1 g in weight. Place 10 ml of conc. HCl in a 4-oz. bottle and fill the bottle to the brim with H$_2$O. Invert this bottle in a pan half filled with water and introduce the weighed magnesium into the bottle. When the reaction has ceased, equalize the levels of the water inside and outside the bottle, and, with the aid of a glass plate, remove the bottle of water and gas to the desk top. Set it upright and, with a glass-marking pencil, mark the level of the water. Then, fill the bottle to the brim with water, carefully recording the amount required. Record this figure as the volume of hydrogen produced. Take the barometer and thermometer readings.

II: DATA

<table>
<thead>
<tr>
<th>Volume of H$_2$ evolved</th>
<th>Volume of H$_2$ at STP</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Weight of H$_2$ evolved...

(1 liter of H$_2$ at STP weighs 0.08987 g)

Weight of H$_2$ evolved by 0.1 g Mg...

Weight of H$_2$ evolved by 1 g atomic weight of magnesium...

Valence of magnesium...
EXPERIMENT 18

DETERMINATION OF THE PERCENTAGE OF COPPER IN COPPER OXIDE

and of THE FORMULA FOR CUPRIC OXIDE

I: EXPERIMENT

Set up the apparatus as in the diagram:

If we know the weight of copper in a given weight of a copper compound we can determine the percent of copper that compound contains. If we have a binary compound, we can calculate also the percent of the other component, and then, knowing the complete percentage composition, we can posit the formula of the compound.

II: DATA

<table>
<thead>
<tr>
<th>Description</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Weight of tube X + copper oxide</td>
<td></td>
</tr>
<tr>
<td>Weight of tube X empty</td>
<td></td>
</tr>
<tr>
<td>Weight of the copper oxide</td>
<td></td>
</tr>
<tr>
<td>Weight of tube X + copper, after heating</td>
<td></td>
</tr>
<tr>
<td>Weight of the copper</td>
<td></td>
</tr>
<tr>
<td>Percent of copper</td>
<td></td>
</tr>
<tr>
<td>Theoretical percent of copper</td>
<td></td>
</tr>
<tr>
<td>Percent error</td>
<td></td>
</tr>
<tr>
<td>Percent of oxygen</td>
<td></td>
</tr>
<tr>
<td>Formula for copper oxide</td>
<td></td>
</tr>
</tbody>
</table>

DO CALCULATIONS HERE
EXPERIMENT 19
TYPES OF CHEMICAL REACTION

I: COMBINATION or SYNTHESIS
1: Using a deflagrating spoon or a metal spatula, heat a bit of sulfur and iron filings together.
2: Using a forceps, heat a small piece of magnesium ribbon in the flame.

II: DECOMPOSITION or ANALYSIS
1: Heat a small pinch of HgO in a small test tube and test the escaping gas. Identify also the substance that collects on the walls of the tube.
2: Heat a few marble chips in a test tube fitted with a type "C" delivery tube extending into a small test tube containing 2 ml Ca(OH)₂.

III: DISPLACEMENT or SINGLE REPLACEMENT
1: Place separate small portions of granulated zinc, iron fillings, and magnesium powder in separate test tubes. Add 5 drops of dil. HCl to each. Identify the escaping gases.

IV: DOUBLE DECOMPOSITION, DOUBLE REPLACEMENT or METATHESIS
1: To 1 ml of dil. HCl in a test tube, add 3 drops of AgNO₃. This reaction is the test for the presence of the chloride ion.
2: To 1 ml of Pb(NO₃)₂ in a test tube, add a few drops of K₂CrO₄. The yellow precipitate is a confirmatory test for the presence of lead.
3: Add 2 or 3 drops of dil. HCl to a crystal of CaCO₃ in a test tube. The fizzing action that results is the test geologists use to identify a carbonate.
4: To 1 ml of Na₂SO₄ in a test tube, add 2 drops of BaCl₂. The white precipitate indicates the presence of a sulfate, provided it is insoluble in acid.

REPORT
These experiments may be reported by defining the type of reaction and placing the equations for the type under the definition. Henceforth, classify all reactions for type as you write them.
EXPERIMENT 20

PREPARATION AND PROPERTIES OF ACIDS AND BASES

I: GENERAL METHODS OF PREPARING ACIDS

1: Salt of the acid desired + sulfuric acid
2: Acid anhydride + water
3: Direct combination (for binary acids only)

II: PREPARATION AND PROPERTIES OF ACIDS

Place a spatulaful of NaCl in a small test tube. Moisten with H₂SO₄ and mix the contents well by stirring or shaking. Exhale cautiously across the mouth of the tube. Bring fumes of NH₃ into contact with the gas escaping from the generator. Test the fumes from the generator with moist litmus paper.

Place a spatulaful of Na₂SO₃ or NaHSO₃ in a test tube and moisten it with H₂SO₄. Carefully POUR the GAS into a test tube half full of water, being careful not to inhale the fumes. Stopper and shake the tube well and then test the resulting solution with litmus paper. Test with phenolphthalein. Pour a few drops of the solution on a crystal of Na₂CO₃ on a glass plate or watch glass and observe the results.

Place a tiny bit of sulfur in a deflagrating spoon or on the tip of your spatula and ignite it in the burner. Hold the burning sulfur over 5 ml of water in a beaker that is covered partly with a watch glass. Remove the unburned sulfur and shake the H₂O and SO₂ fumes until the latter have dissolved. Test the solution as before, with litmus, phenolphthalein and a carbonate.

III: GENERAL METHODS OF PREPARING A BASE

1: Salt of the base desired + a strong base
2: Basic anhydride + water
3: Active metal + water

IV: PREPARATION AND PROPERTIES OF BASES

Take a small strip of magnesium and ignite it by holding it in the flame by the aid of a pair of forceps. DO NOT LOOK DIRECTLY AT THE FLAME. Dissolve the resulting powder in 2 ml of water. Feel this solution; test it with litmus paper and with phenolphthalein.

In a small test tube, dissolve a bit of quicklime in 2 ml H₂O and test this solution as above.

Take a piece of sodium the size of a pin head and place it in a small beaker containing 10 ml H₂O. Cover immediately with a watch glass and, when all action has ceased, test the solution with litmus and phenolphthalein.

Place a pinch of NH₄Cl in a test tube and add 2 ml conc. NaOH.
### Physical properties of acids:

<table>
<thead>
<tr>
<th>C-1</th>
<th>a: density</th>
<th>( \text{H}_2\text{SO}_4 )</th>
<th>HCl</th>
<th>( \text{HNO}_3 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>b: color</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>c: odor</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>d: taste</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>e: solubility</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Chemical properties of acids:

<table>
<thead>
<tr>
<th>C-2</th>
<th>a: stability</th>
<th>( \text{Na}_2\text{CO}_3 + \text{acid} )</th>
<th>( \text{Base} + \text{acid} )</th>
<th>NaOH</th>
<th>Ca(OH)₂</th>
<th>NH₄OH</th>
</tr>
</thead>
<tbody>
<tr>
<td>b: combustibility</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>c: reaction with elements</td>
<td></td>
<td></td>
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<td></td>
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<tr>
<td>d: reaction with compounds</td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>e: reaction to litmus</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>reaction to phenolphthalein</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ionization</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>f: test for identification</td>
<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
</tbody>
</table>

### Physical properties of bases:

<table>
<thead>
<tr>
<th>C-1</th>
<th>a: density</th>
<th>NaOH</th>
<th>Ca(OH)₂</th>
<th>NH₄OH</th>
</tr>
</thead>
<tbody>
<tr>
<td>b: color</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>c: odor</td>
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<td></td>
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</tr>
<tr>
<td>d: taste</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>e: solubility</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>g: feeling</td>
<td></td>
<td></td>
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<td></td>
</tr>
</tbody>
</table>

### Chemical properties of bases:

<table>
<thead>
<tr>
<th>C-2</th>
<th>a: stability</th>
<th>NaOH</th>
<th>Ca(OH)₂</th>
<th>NH₄OH</th>
</tr>
</thead>
<tbody>
<tr>
<td>b: combustibility</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>d: reaction with compounds</td>
<td></td>
<td></td>
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<td></td>
</tr>
<tr>
<td>e: reaction to litmus</td>
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<td></td>
</tr>
<tr>
<td>reaction to phenolphthalein</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ionization</td>
<td></td>
<td></td>
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<td></td>
</tr>
</tbody>
</table>
EXPERIMENT 21

QUANTITATIVE DETERMINATION OF THE ACIDITY
OF VINEGAR

Vinegars are mixtures of water, acids (chiefly acetic), coloring, and other organic substances. We are going to determine what percentage of vinegar is acid. We shall use a measured amount of NaOH of known concentration and, with it, neutralize a known amount of vinegar. Remembering that $NV = N'V'$, we can then calculate the normality of the vinegar. From this we will know the molarity and therefore can calculate the percentage of acid present.

I: PROCEDURE

Make up 250 ml of 0.01N NaOH by weighing out 0.1 g NaOH pellets and dissolving them in sufficient water to make up a total volume of 250 ml. Shake this solution well and transfer some of it to a burette.

By the aid of a pipet, transfer 1 ml of vinegar from the commercial stock bottle into a flask. Dilute it with 9 ml of $H_2O$. Pipet 1 ml of this diluted vinegar into an Erlenmeyer flask and add some water so that the bottom of the flask is well covered. Add 2 drops of phenolphthalein and then let the NaOH enter the solution dropwise from the burette until the first permanent pink color is formed in the flask. (Be sure to rotate the flask constantly to mix the incoming NaOH with the vinegar solution.) Take the reading of the burette and calculate the normality of the vinegar solution. You have now enough information to calculate the normality, molarity, and percentage of vinegar in the original sample of vinegar.

(Note: If there are no burettes available, a pipet may be rigged up with a rubber tube and pinch cock at the top, and this can serve as a burette.)

II: DIAGRAM
DATA SHEET FOR EXPERIMENT 21

Determination of the Acidity of Vinegar

First Trial
Total amount of vinegar used.................................................................
Dilution of vinegar...............ml to...................................................
Final NaOH reading.................................................................
Initial NaOH reading.............................................................
Total NaOH (by difference)..........................................................
Concentration of NaOH in grams per ml............................................

Calculations: Indicate steps used and compare final percentage with that indicated on bottle label.

Second Trial
Total amount of vinegar used.................................................................
Dilution of vinegar...............ml to...................................................
Final NaOH reading.................................................................
Initial NaOH reading.............................................................
Total NaOH (by difference)..........................................................
Concentration of NaOH in grams per ml............................................

Calculations:

Acidity of vinegar present in percent................................................
Acidity of vinegar in moles per liter.................................................
Percent of error...........................................................................

Third Trial: (May be necessary if there is a wide discrepancy in the results of the first two trials.)
EXPERIMENT 22
PREPARATION OF AN INSOLUBLE SALT

I: DISCUSSION

In general, salts may be prepared by one of four major methods which remove interfering substances and permit the recovery of pure salts.

*Direct Combination* is possible for binary salts as then there are no interfering substances.

*Neutralization* of a base by an acid yields a salt and water, which latter substance may be removed by evaporation.

*Volatility* of one of the products removes the contaminant from the field of action. The reaction of a carbonate with an acid is an example of this type of reaction.

*Insolubility* of one of the products also serves to separate a pure salt from the reaction field. This is the method employed in preparing insoluble salts. However, an insoluble salt can be prepared from two soluble salts only, if it is to be pure; otherwise it would contain traces of the insoluble reagent in the product.

II: EXPERIMENT

In this experiment we shall make three insoluble salts.

1: *From the reaction between a salt and an acid.* In a test tube, place 1 ml of Pb(NO₃)₂ solution and add to it 3 drops of dil. HCl.

2: *From a reaction between two salts.* In a test tube place 1 ml Na₂SO₄ solution. Add to it 5 drops of BaCl₂. Centrifuge and isolate the precipitate.

3: *From an oxide and an acid.* In a test tube, dissolve a tiny pinch of PbO in hot, dilute HCl. Dilute the resulting solution with 5 ml of boiling water, centrifuge, and pour off the supernatant liquid.
REPORT FORM FOR EXPERIMENT 22
Preparation of Insoluble Salts

Write the equations for all the reactions.
Write balanced ionic equations for all reactions.
Label the type of reaction (displacement, analysis, etc.)
Tell why each reaction goes to completion.

I: \[ \text{Pb(NO}_3\text{)}_2 + \text{HCl} \rightarrow \]

ionic equation

type reaction

completion?

II: \[ \text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \]

ionic equation

type reaction

completion?

III: \[ \text{PbO} + \text{HCl} \rightarrow \]

ionic equation

type reaction

completion?
DISCUSSION

The four general methods of preparing a salt are:

1: direct combination  
2: neutralization  
3: volatility of one product  
4: insolubility of one product

EXPERIMENT

PREPARATION OF SOLUBLE SALTS

I: DISCUSSION

II: EXPERIMENT

1: Preparation from a metal and an acid. To a tiny piece of zinc in a small test tube, add 2 ml dil. HCl and let the reaction proceed until complete. Centrifuge the remainder and evaporate the supernatant liquid almost to dryness.

2: Preparation from a metallic oxide and an acid. Dissolve 0.5 g of MgO in dil. H₂SO₄. Centrifuge the solution and evaporate the supernatant liquid almost to dryness.

3: Preparation from a carbonate and an acid. To 2 ml Na₂CO₃ solution in a test tube, add about 1 ml dil. HCl and heat. When the effervescence ceases, evaporate the remainder almost to dryness.

4: Preparation by the neutralization of an acid and a base. Add dilute NaOH to 2 ml dil. HCl in a test tube until the resulting solution is neutral to litmus. Evaporate to dryness and taste the residue.

Note: See if you can devise methods to test these salts to prove that they are really what the equations state them to be.
REPORT FORM FOR EXPERIMENT 23

Soluble Salts

Write equations for all reactions.
Write balanced ionic equations for all reactions.
Tell the type of reaction (synthesis, analysis, etc.)
Tell why each reaction goes to completion.

I: Zn + HCl →
ionic reaction
type of reaction?
completion?

II: MgO + H₂SO₄ →
ionic reaction
type of reaction?
completion?

III: Na₂CO₃ + HCl →
ionic reaction
type of reaction?
completion?

IV: NaOH + HCl →
ionic reaction
type of reaction?
completion?

Any additional equations used to prove substances formed to be what were expected may be added here or may be written in the section to which they apply.
I: DISCUSSION

When salts are dissolved in water they ionize, and although water is a very poor conductor of electricity, it too is ionized very slightly. Hence, the ions of the salt can combine with the ions of the water to regenerate small amounts of the acid and base from which the salt was formed. If the original acid was strong and the base combined with it was weak, the salt would produce an acid effect on litmus paper as a result of the regeneration of this strong acid. Similarly, if the original base were strong, the resulting solution would be basic. If both acid and base were equally strong, the salt solution would be neutral.

II: EXPERIMENT

Dissolve the following salts in water and test the resulting solution for acidity with litmus paper or Universal indicator. Report your findings in the following chart and write equations for all reactions (ionic equations).

<table>
<thead>
<tr>
<th>SALT</th>
<th>FORMULA</th>
<th>LITMUS REACTION</th>
<th>ACID FORMED</th>
<th>WEAK OR STRONG</th>
<th>BASE FORMED</th>
<th>WEAK OR STRONG</th>
<th>EQUATION</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na₂CO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₄NO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca(C₂H₃O₂)₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>PbCl₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgSO₄</td>
<td></td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CuSO₄</td>
<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>Al(NO₃)₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>BaCl₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pb(C₂H₃O₂)₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
I: PRINCIPLES OF IONIZATION

By means of the apparatus diagrammed in Fig. 21 below, the relative conducting power of solutions may be determined and therefore their relative degree of effective ionization.

![Apparatus Diagram]

II: EXPERIMENT

With the set-up illustrated above, place 5 drops of each of the following substances on glass plates or watch glasses: H₂SO₄ conc., H₂SO₄ dil., HC₂H₃O₂, NH₄OH conc. and dil., NaOH 10 percent, NaCl, sugar, alcohol, KNO₃.

Successively touch the electrodes to the solutions at about the same distance apart in all cases. Observe whether or not the lamp glows. If it does not, record the substance as a non-conductor. If it does glow, note the comparative brilliancy of the lights. Dip the electrodes in water to wash them between tests.

Record your observations on the following page.

(Note: This experiment may be performed with regular commercial electrolysis apparatus in which a few milliliters of the solutions are placed in small beakers and tested. Evaporating dishes or crucibles may be substituted for beakers if desired.)
REPORT FORM FOR EXPERIMENT 25

Conductivity

<table>
<thead>
<tr>
<th>SUBSTANCE</th>
<th>GOOD</th>
<th>POOR</th>
<th>FAIR</th>
<th>NON</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{H}_2\text{SO}_4$, conc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{H}_2\text{SO}_4$, dil.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HNOS}_3$, conc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HNOS}_3$, dil.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HC}_2\text{H}_3\text{O}_2$, conc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HC}_2\text{H}_3\text{O}_2$, dil.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HCl}$, conc.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HCl}$, dil.</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{NH}_4\text{OH}$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{NaOH}$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{NaCl}$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sugar</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Alcohol</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{KNO}_3$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Write equations for the ionization of those solutions which are good conductors.
DISCUSSION

As you know, any metal in the EMF Series will replace the ion of a metal lower in the series; hence, aluminum will replace copper ions from solutions of copper salts. So, if we have a solution of copper sulfate and dip into it a piece of aluminum, we shall expect the aluminum to become coated with copper.

II: EXPERIMENT

Thoroughly clean a piece of aluminum metal and place it in a solution consisting of 0.3 g CuSO\textsubscript{4} and an equal amount of cream of tartar and Na\textsubscript{2}CO\textsubscript{3}, with 10 g of H\textsubscript{2}O. When the aluminum is all copper-coated, remove it from the solution, wash, and dry it.

A variation of this may be done as a project or demonstration by placing clean strips of copper sheet in the form of a tree skeleton and immersing the whole in a beaker of AgNO\textsubscript{3}. In time, the “tree” will grow into a myriad of silver crystals.

Zinc may be plated as was aluminum, in the first part of the experiment, but with the copper solution made up as follows: 0.3 g NH\textsubscript{4}OH, 5.0 g H\textsubscript{2}O and 0.3 g CuSO\textsubscript{4}.

III: ELECTROPLATING

Connect a strip of copper to a positive terminal and a carbon rod to the negative terminal of an electric current source. The wires must be connected to a direct-current source (DC, dry cell or storage battery). Immerse both electrodes in a CuSO\textsubscript{4} solution, being sure not to touch them to each other. Allow them to remain thus fixed until there is a definite change in the appearance of the carbon. Reverse the wires until the copper has all been removed and replated on the copper strip. Return this copper sulfate solution to the stock reagent bottle and dry the equipment.
EXPERIMENT 27
HYDROGEN AS A REDUCING AGENT

I: EXPERIMENT

Place a few pieces of mossy zinc in a 4-in. test tube and connect a right angle delivery tube. To the other end of this delivery tube, attach, by means of rubber tubing, a 4-in. section of glass tubing. Plug both ends of this section loosely with cotton after you have spread a bit of copper oxide, wire form, in the length of the tube. Now, by means of another section of rubber tubing, connect a capillary to the farther end of the set-up. Pour 2 ml dil. HCl into the generator tube and place the burner under the tube containing the CuO. Continue to heat this tube until the contents change color noticeably.

II: QUANTITATIVE METHOD

This method can be rendered quantitative by weighing the tube of copper oxide before and after heating. The weight lost is the weight of O₂ driven off. From this, the amount of H₂ used can be calculated and a quantitative determination of the amount of copper in the copper oxide can be made, as well as the amount of H₂ sent over by the acid, though this will not be exact, as the excess will merely go through the system. If, however, a weighed amount of CaCl₂ in a tube is placed in the line beyond the CuO tube, the increase in weight of this tube at the end of the experiment will be the amount of water formed. Your teacher can tell you what variations of this experiment he wishes you to perform.

III: DIAGRAM

![Diagram of experimental setup](Fig. 22)
EXPERIMENT 28
PREPARATION AND PROPERTIES OF CHLORINE

I: PREPARATION

Set up the apparatus for the collection of a gas heavier than air and soluble in water (gas set-up "C" Fig. 17) and place about a gram of MnO₂ in the generator. To this, add 3 ml conc. HCl and collect 6 vials of gas by the displacement of air. It may be necessary to heat the generator slightly. As each vial is filled, stopper it immediately. The color of the gas is visible and will serve to indicate when the vials are filled. Take another vial and place 5 ml H₂O in it and allow the delivery tube to rest under the surface of the water while the gas bubbles in.

II: PROPERTIES

Physical
a: Calculate the density of chlorine gas. [C-1a]
b: Note the color, the odor and the solubility. [C-1b, C-1c] DO NOT TASTE

Chemical
a: Heat a spatulaful of powdered antimony and sprinkle into a vial of gas. [C-2c]
b: Repeat with heated iron filings. [C-2c]
c: Repeat with heated copper filings. [C-2c]
d: Test the combustibility with a burning splint. [C-2b]
e: Moisten a piece of colored cotton and insert it in a flask or vial of Cl₂ gas. [C-2e]
f: To 2 ml of KBr, add 1 ml Cl₂-H₂O, obtained by bubbling Cl₂ into water. Add 1 ml CCl₄ and shake. Observe the color.
g: Repeat the above using a solution of KI.
h: Look up the test for Cl₂ and apply it to the gas.

III: PROBLEMS

How much chlorine can be obtained by the complete oxidation of 5 g of HCl?

If KClO₃ and HCl are used as a source of Cl₂, what weight of HCl will be necessary to produce 50 ml of Cl₂ gas? If we start with 12N HCl, what volume must we use?
OUTLINE REPORT FORM FOR CHLORINE

A: OCCURRENCE
   1: Places found free
   2: Places found combined

B: PREPARATION

   Laboratory
   1: MnO₂ + HCl →
   2: KMnO₄ + HCl →
   3: KClO₃ + HCl →
   4: PbO₂ + HCl →

   Commercial

C: PROPERTIES

   1: Physical
      a:
      b:
      c:
      d:
      e:
      f:

   2: Chemical
      a:
      b:
      c: Fe + Cl₂ →
         Cu + Cl₂ →
         Sb + Cl₂ →
      d: Cl₂ + H₂O →
         Cl₂ + KBr →
         Cl₂ + KI →
      e: Powerful oxidizing agent
D: ATOMIC STRUCTURE
1: atomic diagram 2: molecular diagram 3: ionic diagram

Placement in Periodic Table
Metal or non-metal
Electron donor or receiver

E: USES

F: OTHER IMPORTANT EQUATIONS
\[
\begin{align*}
H_2 & + Cl_2 \rightarrow \\
As & + Cl_2 \rightarrow \\
Cl_2 & + H_2O \rightarrow \\
Na & + Cl_2 \rightarrow \\
NaCl & + H_2SO_4 + MnO_2 \rightarrow
\end{align*}
\]
EXPERIMENT 29

PREPARATION AND PROPERTIES OF BROMINE

What quantity of bromine can be prepared by the complete oxidation of 5 g of KBr?

How much KBr is needed to prepare 1 ml of bromine?

I: LABORATORY PROCEDURE

CAUTION: BROMINE IS EXTREMELY CORROSIVE TO THE SKIN IN THE LIQUID AND VAPOR PHASES. AVOID ALL DIRECT CONTACT WITH IT. IN CASE OF ACCIDENT, WASH THE AREA CONTAMINATED AND THEN RUB IT WELL WITH ALCOHOL, FOLLOWED BY GLYCERINE.

Using gas set-up “C” (Fig. 17), place 0.5 g KBr and half as much MnO₂ in the generator. Add 2 ml conc. H₂SO₄ and let the delivery tube dip just below the surface of water in the collecting vial. Discontinue heating and simultaneously withdraw the delivery tube from the water when the generator ceases to be filled with brown fumes.

II: PROPERTIES

1: Physical
   a: Determine the color, odor, solubility and density of bromine, but do NOT directly inhale the fumes. [C-1a, C-1b, C-1c]
   b: The method of collection makes clear the answer to C-1e and C-1f.
   c: Add 1 drop of bromine water to a drop of CCl₄ and shake well. [C-1e]

2: Chemical
   a: Place a drop of bromine water on litmus paper and watch the changes. [C-2e]
   b: Dissolve a speck of KI in water; add 1 drop of bromine water and 4 drops of CCl₄. Shake well and note the color. [C-2d]
   c: Dissolve a pin-head crystal of KBr in 1 ml of water; add 2 drops of chlorine water and 2 drops of CCl₄. Shake well and note the color. [C-2c]

III: REACTION

If you have time, see if you can determine the identity of the substances remaining in the generator. Look up the tests necessary to confirm the presence of the ions indicated in your general equation for the preparation of bromine.
OUTLINE REPORT FORM FOR BROMINE

A: OCCURRENCE
1: Places found free
2: Places found combined
   Method of extraction

B: PREPARATION
1: Laboratory
2: Commercial

C: PROPERTIES
1: Physical
   a:
   b:
   c:
   e:
   f:
   g:

2: Chemical
   a:
   b:
   c:
   d: \( \text{Cl}_2 + \text{KBr} \rightarrow \)
   \( \text{Br}_2 + \text{KI} \rightarrow \)
   \( \text{Br}_2 + \text{Litmus paper} \rightarrow \)
   e:
   f:
D: ATOMIC STRUCTURE

1: atomic diagram    2: molecular diagram    3: ionic diagram

Placement in Periodic Table
Metal or non-metal
Electron donor or receiver

E: USES

F: OTHER EQUATIONS

\[ \text{NaBr} + \text{AgNO}_3 \rightarrow \]
\[ \text{C}_2\text{H}_2 + \text{Br}_2 \rightarrow \]
EXPERIMENT 30

PREPARATION AND PROPERTIES OF IODINE

I: LABORATORY PREPARATION
Place about a spatulaful of \( \text{MnO}_2 \) in a crucible and add half as much \( \text{KI} \). To this add 2 drops of conc. \( \text{H}_2\text{SO}_4 \) and cover with a watch glass containing a few drops of cold water or a crystal of ice. Heat VERY gently until the vapors have subsided. Remove the flame, let the crucible cool, and then carefully remove the watch glass, being sure that none of the water contained in it runs over onto the under side. Examine the crystals adhering to the glass.

Fig. 23

II: PROPERTIES

1: Physical
   a: Determine the color, odor and shape of the crystals. DO NOT TASTE. [C-1a, C-1b, C-1c, C-1g]
   b: Determine the relative solubility of \( \text{I}_2 \) in \( \text{H}_2\text{O}, \text{C}_2\text{H}_5\text{OH}, \text{CCl}_4 \).
   c: Place a speck of \( \text{KI} \) in half a test tube of \( \text{H}_2\text{O} \), and try to dissolve a bit of \( \text{I}_2 \) in this solution. [C-1e]
   d: Note that the Iodine changed directly from the solid to the vapor state [C-1f]

2: Chemical
   a: Place 1 ml \( \text{NaOH} \) in a test tube and add a crystal of \( \text{I}_2 \). [C-2d]
   b: Test the foregoing solution with litmus and write a redox reaction that accounts for the results.
   c: In a small test tube, place a speck of \( \text{KI} \). Dissolve it in water, add 2 drops of chlorine water, 2 drops \( \text{CCl}_4 \) and shake. [C-2c]
   d: Repeat the above using bromine water in place of chlorine water.
   e: Boil 1 ml of water and add to it a tiny speck of starch. (Enough to cover the tip of a spatula.) Take \( \frac{1}{2} \) ml of this solution in a test tube and add 2 drops of a solution made by dissolving \( \text{I}_2 \) in \( \text{KI} \) solution. Note the formation of a dark blue color. [C-2f]

III: PROBLEM
If a certain seaweed is 12 percent iodine in the form of \( \text{KI} \), how much seaweed must be gathered to produce 10 g of iodine?
OUTLINE REPORT FORM FOR IODINE

A: OCCURRENCE
   1: Places found free
   2: Places found combined
      Method of extraction

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      d:
      e:
      f:
   2: Chemical
      a:
      b:
      d: \( \text{Cl}_2 + \text{KI} \rightarrow \text{Br}_2 + \text{KI} \rightarrow \text{NaOH} + \text{I}_2 \rightarrow \)
      e:
      f:

D: STRUCTURE
   1: atomic diagram  2: molecular diagram  3: ionic diagram

E: USES

F: OTHER IMPORTANT EQUATIONS
   \( \text{NaI} + \text{AgNO}_3 \rightarrow \)
EXPERIMENT 31
PREPARATION AND PROPERTIES OF HYDROCHLORIC ACID AND HYDROGEN CHLORIDE

I: LABORATORY PREPARATION

Using gas set-up "C" (Fig. 17), place about a gram of NaCl in the generator and add 2-3 ml dil. H₂SO₄. Collect 3 vials of the gas by the displacement of air. The tubes will fume at the mouth when filled if the day is damp; otherwise, blow across the top of the collecting vial, and, if it is full, the fumes will be visible. Fill another collecting vial with water and bubble some of the gas into it to prepare hydrochloric acid.

II: PROPERTIES

1: Physical
   a: Calculate the density of the gas. How much salt would be needed to prepare 25 ml of HCl gas? If the collection were carried on at 758 mm pressure and 32°C, what volume would this gas occupy at standard conditions?
   b: Note the color and odor of the gas. [C-1b, C-1c]
   c: Blow across the vial if you have not already done so in the collecting process. This resultant fuming is a rapid test for the detection of a hydrohalide.
   d: Invert a vial of the gas in a beaker of H₂O. [C-1e] Taste the solution that you then find in the tube. [C-1d]

2: Chemical
   a: Touch a drop of this solution, or of the solution in the last collecting vial of HCl + H₂O, to litmus paper. [C-2e]
   b: Test the combustibility of a tube of the gas. [C-2b]
   c: Carefully smell a bottle of HCl from the stockshelf. The odor of HCl gas is evidence of the lack of stability of the solution. [C-2a]
   d: Into a ml of the HCl solution, put a few Mg turnings and test the gas evolved, by the use of a burning splint. [C-2c]
   e: Place a marble chip in a test tube or on a glass plate and add to it 2 drops of your acid. [C-2d]
   f: If you have done the experiment on conductivity, try out the conductivity of your acid. [C-2e]

III: TEST FOR THE CHLORIDE ION (Cl⁻)

To a milliliter of any chloride solution, add 2 drops of dil. HNO₃, and 2 drops of AgNO₃. A white, curdy precipitate is indication of the presence of a chloride. This precipitate will darken upon exposure to direct sunlight.

If NH₄OH is added until the precipitate first obtained is dissolved, the silver ammonia complex will form. The readdition of AgNO₃ will reprecipitate the AgCl.
OUTLINE REPORT FORM FOR
HYDROCHLORIC ACID and HYDROGEN CHLORIDE

A: OCCURRENCE

B: PREPARATION
1: Laboratory
2: Commercial

C: PROPERTIES
1: Physical
   a: 
   b: 
   c: 
   d: 
   e: 
   f: 
2: Chemical
   a: 
   b: 
   c: 
   d: 
   e: 
   f: 

D: STRUCTURE
1: Molecular diagram

E: USES

F: OTHER IMPORTANT EQUATIONS
   AgCl + NH₄OH → 
   NaCl + H₂SO₄ → 
   Fe + HCl → 
   HCl + NaOH → 
   HCl + ZnO → 
   HCl + MnO₂ →
EXPERIMENT 32
PREPARATION AND PROPERTIES OF SULFUR DIOXIDE

I: LABORATORY PREPARATION

Sulfur dioxide may be prepared by the burning of sulfur, the heating of a bisulfite, or the reduction of sulfuric acid with copper turnings. However, for laboratory quantities, the best results are obtained by reacting an acid with a bisulfite.

Into one empty collecting vial, place a piece of moist, neutral litmus paper.
Into another vial, place a few pieces of moist green grass or fern.
Into a third, place moistened light color cloth or rose petals.
Into two more vials place small slices of peeled apple and cork these vials until you are ready to use them.

Using gas set-up type “C” (Fig. 17), place about 1 g of NaHSO₃ in the generator and add 2 ml dil. HCl. Insert the delivery tube and pass the gas into the first of the vials prepared. When this is filled, stopper it and fill the other vials in order, and lastly, fill one of the tubes containing the apple slices.

II: PROPERTIES

1: Physical
   a: Observe the color, odor and taste of the gas. [C-1b, C-1c, C-1d]
   b: Calculate the density of the gas. [C-1a]
   c: Determine the solubility of the gas. [C-1e]

2: Chemical
   a: Observe the reactions that have occurred in the vials containing the various substances. [C-2d, C-2e]
   b: Test the gas with moist litmus paper. [C-2e]
   c: Place dil. KMnO₄ in a test tube and add a solution made by dissolving SO₂ in H₂O. Observe the results obtained by shaking these solutions. [C-2f]
   d: Leave the apple slices until the next lab. period and then compare the slices in the vial containing SO₂ with those in the vial containing air.

III: PROBLEM

If you prepared 40 ml of SO₂, how much bisulfite was necessary? Did you use this amount? If not, why not?
OUTLINE REPORT FORM FOR SULFUR DIOXIDE

A: OCCURRENCE

B: PREPARATION
1: Laboratory
   a: Burning sulfur
   b: Heating bisulfite
   c: Reducing H₂SO₄
   d: Acid on a bisulfite
2: Commercial

C: PROPERTIES
1: Physical
   a:
   b:
   c:
   d:
   e:
   f:
2: Chemical
   a:
   b:
   d: SO₂ + H₂O →
      SO₂ + KMnO₄ →
   e:
   f:

D: STRUCTURE
1: Molecular diagram

E: USES

F: OTHER EQUATIONS
   SO₂ + H₂O →
EXPERIMENT 33
PREPARATION AND PROPERTIES OF HYDROGEN SULFIDE

I: PREPARATION

Gas set-up type “C” (Fig. 17) may be used, but the poisonous and evil-smelling gas is quite likely to escape and fill the room in a short time. Hence, it is considered desirable to use the following method.

Connect two CaCl₂ tubes by means of about 6 in. of rubber tubing, and into one tube, put the type “C” delivery tube. On the free end of this tube attach a dropper tip by means of a section of rubber tubing, which, in turn, is clamped by a pinch clamp.

Place a couple of lumps of FeS in tube “B” and fill tube “A” with dil. HCl until the FeS is covered. (Have the pinch clamp open during this step.) Close the clamp and the generator is ready to use.

Determine the density of the H₂S gas and decide whether to collect it by the upward or downward displacement of air. Collect two tubes of H₂S and cork them well. (A convenient method of telling when the vial is full is to place a piece of moist lead acetate paper at the mouth of the vial. It will turn black when the gas begins to escape from the filled vial.)

II: PROPERTIES

a: By now you probably know the color and odor and H₂S as well as its density [C-1a, C-1b, C-1c]

b: Add 3 drops of H₂O to 1 vial of the gas and recork it immediately. Shake it well and then taste the resulting solution. [C-1d] Test this solution with litmus paper. [C-2e] The name given to this is hydrosulfuric acid.

c: Test the combustibility of one vial of the gas. [C-2b]

d: Place 1 ml Pb(NO₃)₂ in a test tube. Bring the tube under the end of the delivery tube of the generator and make sure the solution covers the tip. Turn on the generator and note the change in the color of the solution. Repeat this step with solutions containing cadmium ions, arsenic ions and antimony ions. [C-2d]
OUTLINE REPORT FORM FOR HYDROGEN SULFIDE

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      d:
      e:
   2: Chemical
      a:
      b:
      d: \( \text{H}_2\text{S} + \text{Pb(C}_3\text{H}_3\text{O}_2)_2 \rightarrow \) 
      \( \text{H}_2\text{S} + \text{Pb(NO}_3)_2 \rightarrow \) 
      \( \text{H}_2\text{S} + \text{Cd(NO}_3)_2 \rightarrow \) 
      \( \text{H}_2\text{S} + \text{Zn(NO}_3)_2 \rightarrow \) 
      \( \text{H}_2\text{S} + \text{AsCl}_3 \rightarrow \) 
      \( \text{H}_2\text{S} + \text{SbCl}_3 \rightarrow \) 
      e:
      f:

D: STRUCTURE

E: USES

F: OTHER IMPORTANT EQUATIONS
   \( \text{H}_2\text{S} + \text{BiCl}_3 \rightarrow \) 
   \( \text{H}_2\text{S} + \text{NaCl} \rightarrow \) 
   \( \text{H}_2\text{S} + \text{CuSO}_4 \rightarrow \) 
   \( \text{H}_2\text{S} + \text{KMnO}_4 \rightarrow \) 
   \( \text{H}_2\text{S} + \text{K}_2\text{CrO}_4 \rightarrow \) 
   \( \text{H}_2\text{S} + \text{O}_2 \rightarrow \)
EXPERIMENT 34
PREPARATION AND PROPERTIES OF AMMONIA AND AMMONIUM HYDROXIDE

I: PREPARATION

Using gas set-up type "B", mix 1 g of NH₄Cl with an equal amount of dry Ca(OH)₂ and place them in the generator tube. Heat the mixture gently and collect 3 vials of gas. A piece of blue litmus paper moistened and hung at the mount of the collecting vial will turn color when the vial is filled with gas. When the 3 vials have been collected and stoppered, turn the delivery tube downward and bubble the gas into a vial of water for 2 minutes.

II: PROPERTIES

1: Physical
   a: Calculate the density of NH₃ gas. [C-1a]
   b: Observe the color, odor and taste of ammonia. [C-1b, C-1c, C-1d]
   c: Invert a vial of ammonia in a small beaker of water. [C-1e]

2: Chemical
   a: Test the combustibility of NH₃ by inserting a burning splint into a vial of the gas. [C-2b]
   b: Place an empty vial over the mouth of a bottle of conc. HCl for a minute; withdraw it and place it mouth to mouth with a vial of NH₃. [C-2d]
   c: Test the vial into which NH₃ bubbled, making a solution of NH₃ in H₂O. Test this solution with litmus paper and with phenolphthalein. Also note its color, odor and taste. [C-1b, c, d]
   d: Rub a bit of this NH₄OH solution between your fingers. [C-2e]
   e: Test the conductivity of this solution. [C-2e]

III: TEST FOR THE NH₄⁺ ION

Place about half a gram of an ammonium salt in a test tube and add 5 drops of conc. NaOH. Heat the tube gently while holding a piece of moist red litmus paper near the mouth of the tube. The change in color of this paper, plus the odor characteristic of NH₃ are confirmatory tests for the presence of the ammonium ion.

IV: PROBLEM

Wanted: 1 liter of 0.5N NH₄OH. How much NH₄Cl must be used to generate sufficient gas to make up this solution?
REPORT FORM FOR
AMMONIA and AMMONIUM HYDROXIDE

A: OCCURRENCE

B: PREPARATION
1: Laboratory
2: Commercial
   a: Haber:
   b: Cyanamide:
   d: Destructive distillation of soft coal:

C: PROPERTIES
1: Physical
   a:
   b:
   c:
   d:
   e:
   f:
2: Chemical
   a:
   b:
   d: \( \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{HCl} \rightarrow \)
   e:
   f:

D: STRUCTURE
1: Ammonia molecule

E: USES

F: OTHER EQUATIONS
\( \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow \)
\( \text{NH}_4\text{OH} + \text{HNO}_3 \rightarrow \)
PREPARATION AND PROPERTIES OF NITRIC ACID

I: PREPARATION

The general method of preparation for all acids is employed in the preparation of HNO₃ in the laboratory, viz., the reaction between a salt of nitric acid and high boiling, less volatile sulfuric acid. Because of the corrosive qualities of nitric acid in the vapor state we will not collect the acid by the distillation methods commonly employed.

To a spatulaful of NaNO₃ in a crucible, add 5 drops of conc. H₂SO₄ and cover with a watch glass or crucible lid. Now insert a piece of white wool yarn so that it hangs from the cover, and heat the crucible gently. A change will occur in the yarn, which change is indicative of the presence of HNO₃ and a protein. It is called the xanthropotaic test for nitric acid or for a protein. [C-2f]

II: PROPERTIES

For determining these properties, use the HNO₃ found in the reagent bottles.

a: Color, odor, and taste may be discovered readily. [C-1b, c, d]

b: Add a few drops of very dilute HNO₃ to a small piece of zinc in a test tube. [C-2c]

c: Add a drop of conc. HNO₃ to a copper turning in a test tube. [C-2c]

III: TEST FOR IDENTIFICATION OF THE NO₃⁻ ION

Place 10 drops of a solution of a nitrate in a small test tube and add 10 drops of freshly prepared ferrous ammonium sulfate (or ferrous sulfate). With the tip of the dropper touching the side of the test tube, add 10 drops of conc. H₂SO₄ cautiously so that it forms a separate layer on the bottom of the tube. The formation of a brown ring at the interface is a positive test for a nitrate. [C-2f]

IV: PROBLEM

How much HNO₃ must be used to make up 5 liters of a 20 percent solution? 500 ml of a 0.1N solution? How much 6M KOH would be required to neutralize 20 ml of this latter solution?
OUTLINE REPORT FORM FOR NITRIC ACID

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   (Note: The laboratory preparation is commonly carried out in a retort so that no cork or rubber connections are necessary. The diagram for such a preparation looks like this:

   2: Commercial
      a: Ostwald
      b: Arc Process

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      d:
      e:
   2: Chemical
      a:
      b:
      c: \( \text{HNO}_3 + \text{Zn} \rightarrow \text{conc. HNO}_3 + \text{Cu} \rightarrow \text{dil. HNO}_3 + \text{Cu} \rightarrow \)
      d: \( \text{HNO}_3 + \text{protein} \rightarrow \text{HNO}_3 + \text{NaOH} \rightarrow \)
      e:
      f:

E: USES

F: OTHER IMPORTANT EQUATIONS
   \( \text{NH}_3 + \text{HNO}_3 \rightarrow \)
   \( \text{CaO} + \text{HNO}_3 \rightarrow \)
   \( \text{Zn} + \text{very dil. HNO}_3 \rightarrow \)
   \( \text{N}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \)
EXPERIMENT 36

PREPARATION AND PROPERTIES OF THE
OXIDES OF NITROGEN

I: NITRIC OXIDE (NO)

Set up a gas generator, type “A” (Fig. 15), and place about a gram of copper turnings in the tube. Add 3 ml HNO₃ (diluted 1:1) and stopper the test tube with the stopper attached to the delivery tube which should be well under the water in the collecting trough. Let the first few bubbles of gas escape and then collect 4 vials of nitric oxide. (Note: The gas generated first is mostly NO₂ since the NO united with the air in the generator. However, NO₂ is water soluble, whereas NO is not.)

Test the combustibility of NO in one of the vials. Calculate the density and record all the physical properties you can. If you think the gas has any chemical reactivity you can test, use the remaining vials of gas for these experiments.

II: NITROGEN DIOXIDE (NO₂)

Remove the stopper from a vial of NO collected above and touch a lighted splint to the resulting brown fumes.

Uncork another vial and let it oxidize a moment before adding a few drops of H₂O. Stopper and shake the gas and water and test the resulting solution with litmus paper. (Note: Acid anhydrides are oxides of non-metals that produce acids upon the addition of H₂O.)

Test this solution of NO₂ and H₂O for the nitrate ion by adding fresh ferrous sulfate and conc. H₂SO₄.

Report the physical properties of NO₂.

III: NITROUS OXIDE (N₂O)

Again using gas set-up “A” (Fig. 15), place a gram of NH₄NO₃ in the generator. Heat the tube gently and collect a vial of gas. The success of this experiment depends on the very gentle heating of the solid. Rapid heating will result in many oxidation products and will not give pure N₂O as product.

Report the physical properties of this gas and test its combustibility with a lighted splint.

If you think any of these gases are oxidizing or reducing agents, devise tests to prove that they are.
OUTLINE REPORT. FORM FOR NITRIC OXIDE

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a: 
      b: 
      c: 
      d: 
      e: 
      f: 
   2: Chemical
      a: 
      b: 
      e: 
      f: 

E: USES

F: OTHER EQUATIONS
   \(N_2 + O_2 \rightarrow\)
   \(NH_3 + O_2 \rightarrow\)
   \(Cu + \text{dil. HNO}_3 \rightarrow\)
   \(NO_2 + \text{HNO}_3 \rightarrow\)
   \(Zn + \text{dil. HNO}_3 \rightarrow\)
OUTLINE REPORT FORM FOR NITROGEN DIOXIDE

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      d:
      e:
      f:
   2: Chemical
      a:
      b:
      e:
      f:

E: USES

F: OTHER EQUATIONS
   NO + O₂ →
   Cu + conc. HNO₃ →
OUTLINE REPORT FORM FOR NITROUS OXIDE

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      d:
      e:
      f:
   2: Chemical
      a:
      b:
      c:
      f:

E: USES

F: OTHER EQUATIONS

DIAGRAM OF LAB. PREP.
EXPERIMENT 37
PROPERTIES OF SULFURIC ACID

I: PHYSICAL PROPERTIES
   a: Determine the density of sulfuric acid by carefully weighing 1 ml of it. The figure thus obtained will be the density.
   b: Observe the color, odor, viscosity. [C-1b, C-1c, C-1g]
   DO NOT TASTE.

II: CHEMICAL PROPERTIES
   a: Place 1 ml H₂O in a test tube and add 2 drops of acid. Feel the bottom of the tube. (Consult your text book on "heat of solution" and the hydrates of H₂SO₄)
   b: H₂SO₄ in the concentrated form is a strong dehydrating agent and removes water from carbohydrates leaving only the carbon of the original compound. Dip a wooden splint into some of the acid and examine it after a few minutes.
   c: To a spoonful of sugar on a glass plate, add a few drops of concentrated H₂SO₄ and again observe the evidence of reaction.
   d: Dilute H₂SO₄, on the other hand, behaves as a normal acid and releases H₂ when treated with a metal such as zinc. To 2 ml dil. H₂SO₄ in a small test tube, add a piece of mossy zinc and test the gas evolved by means of a burning splint.
   e: Repeat the procedure with conc. H₂SO₄. In this case the regular acid properties of the H₂SO₄ give way to its oxidizing properties and the gas evolved is not H₂ but SO₂.

III: TEST FOR THE SULFATE ION (SO₄²⁻)
   Add 3 drops of BaCl₂ to 1 ml of any sulfate solution in a test tube. Centrifuge and add 5 drops of HCl to the precipitate. If it is a sulfate, the precipitate will remain. In case there are interfering ions like CO₃⁻² or PO₄⁻³⁻, the HCl will dissolve the barium salt of either.
OUTLINE REPORT FORM FOR SULFURIC ACID

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
   2: Commercial
      a: Lead Chamber
      b: Contact Process

C: PROPERTIES
   1: Physical
      a:
      b:
      c:
      g: hygroscopic
   2: Chemical
      a:
      b:
      c: $\text{H}_2\text{SO}_4 + \text{Fe} \rightarrow$
      dil. $\text{H}_2\text{SO}_4 + \text{Zn} \rightarrow$
      conc. $\text{H}_2\text{SO}_4 + \text{Zn} \rightarrow$
      d: $\text{H}_2\text{SO}_4 + \text{sugar} \rightarrow$
      $\text{H}_2\text{SO}_4 + \text{wood} \rightarrow$
      e:
      f:

E: USES

F: OTHER EQUATIONS
   $\text{H}_2\text{SO}_4 + \text{NaNO}_3 \rightarrow$
   $\text{H}_2\text{SO}_4 + \text{NaCl} \rightarrow$
   $\text{H}_2\text{SO}_4 + \text{Na}_2\text{SO}_3 \rightarrow$
   $\text{H}_2\text{SO}_4$ ionization $\rightarrow$
   2nd step $\rightarrow$
EXPERIMENT 38
CARBON — SOURCES AND FORMS

I: SOURCES OF CARBON

a: Take small portions of any or all of the following and heat them in test tubes until all changes stop: bread, sugar, meat, wood, coal, starch.

b: Shut off the air supply of the burner and light the gas. Hold a glass rod in the luminous flame.

II: FORMS OF CARBON

a: GRAPHITE. Take some crystals of graphite from a bottle on the side shelf and examine them under a hand lens. Try to ignite them. Rub your finger on the graphite and try to write your name with it.

b: AMORPHOUS CARBON. Lampblack: Burn about 5 drops of oil in a crucible. When the oil has completely charred, add another drop and stir the soot into the oil with a match stick. When it is thoroughly mixed, try to write with this substance and after the writing has dried, see if it is removed by water.

Boneblack: Dissolve a gram of brown sugar in 2 ml of water and divide the solution in halves. To one half, add a spatulaful of boneblack, stir well, and heat gently. Centrifuge and pipet off the supernatant liquid into another table. To this add more boneblack and treat as before. Repeat until the liquid resulting is quite clear. Set the liquid aside on a watch glass to evaporate. The purpose of separating the liquid at the outset was to have a standard to compare your color changes with as you proceeded.

Charcoal is another amorphous form of carbon and may be used for decolorization as was boneblack.

c: DIAMOND, the hard, crystalline form, is not used in this course for experimentation.
OUTLINE REPORT FORM FOR CARBON

A: OCCURRENCE
   1: Place and state
   2: Methods of extraction

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a: 
      b: 
      c: 
      d: 
      e: 
      f: 
      g: 
   2: Chemical
      a: 
      b: 
      c: 
      d: 
      e: 
      f: 

D: STRUCTURE
   1: Crystal form

E: USES

F: OTHER EQUATIONS
   C  + O₂ → Fe₂O₃ + C →
   C  + insufficient O₂ → H₂O + C →
   C  + S → CO₂ + C →
   ZnO + C → PbO + C →
EXPERIMENT 39
OXIDES OF CARBON

I: CARBON MONOXIDE

1: Preparation. Fix up generator type “A” (Fig. 15) and place 10 drops of conc. H₂SO₄ and 1 ml formic acid in the tube. Heat very gently and collect 3 vials of CO by the displacement of H₂O. [B-1]

2: Properties. Calculate the density of the gas and observe the other visible properties. Insert a burning splint into the vial of CO to determine its relation to combustion. [C-2b] Light the gas in another vial and immediately add 5 drops of limewater, stopper and shake. [C-2f]

II: CARBON DIOXIDE

How many grams of CaCO₃ are needed to produce 25 ml CO₂?

1: Preparation. Using gas generator type “A”, place 2 g of marble chips in the tube and add 3 ml dil. HCl. Collect 4 vials of gas by water displacement.

2: Properties. Calculate the density of the gas. [C-1a]

Determine the relation to combustion by inserting a lighted splint into a tube of CO₂. [C-2b]

Into another vial of gas, quickly drop 5 drops of limewater and stopper immediately. Shake well and note the change in the limewater. [C-2f]

Into another vial pipet 5 drops of H₂O and, using the thumb as a stopper, shake the solution vigorously and test the mixture with litmus. [C-2d]

Half-fill a test tube with some soft drink and gently heat it. Observe the results. Or, place the soft drink in a test tube fitted with a type “C” delivery tube leading into a tube of Ca(OH)₂ and then gently heat the tube containing the soft drink. Observe the changes in the Ca(OH)₂. [C-1e, C-1f]
A: OCCURRENCE
   1: Coal gas
   2: Water gas
   3: Producer gas

B: PREPARATION
   1: Laboratory
   2: Commercial

C: PROPERTIES
   1: Physical
      a: 
      b: 
      c: 
      e: 
   2: Chemical
      a: 
      b: 
      e: 
      f: 

D: STRUCTURE
   1: Molecular diagram
      Type of valency

E: USES

F: OTHER EQUATIONS
OUTLINE REPORT FORM FOR CARBON DIOXIDE

A: OCCURRENCE

B: PREPARATION
   1: Laboratory
      a: Burning C
      b: Burning carbonaceous matter
      c: Heating a bicarbonate
      d: Acid on a carbonate
   2: Commercial

C: PROPERTIES
   1: Physical
      a: 
      b: 
      c: 
      d: 
      e: 
      f: 
   2: Chemical
      a: 
      b: 
      d: \( \text{CO}_2 + \text{Ca(OH)}_2 \rightarrow \text{CO}_3^2 + \text{H}_2\text{O} \rightarrow \)
      e: acid anhydride
      f: 

D: STRUCTURE
   1: Molecular diagram

E: USES

F: OTHER EQUATIONS
   \( \text{CH}_4 + \text{O}_2 \rightarrow \)
   \( \text{C}_2\text{H}_2 + \text{O}_2 \rightarrow \)
   \( \text{C}_6\text{H}_6 + \text{O}_2 \rightarrow \)
   \( \text{CO} + \text{O}_2 \rightarrow \)
   \( \text{C}_6\text{H}_{12}\text{O}_6 \) fermenting \( \rightarrow \)
   \( \text{Ca(HCO}_3)_2 \rightarrow \)
   \( \text{H}_2\text{SO}_4 + \text{NaHCO}_3 \rightarrow \)
I: COMPONENTS OF THE ATMOSPHERE

1: Nitrogen. Fill a small pan or an evaporating dish, with water to a depth of about 2 cm. Insert a used match through a piece of cardboard about an inch square and then dip the top of the match in red phosphorus and make certain that some of the phosphorus adheres to the top. Immerse the cardboard in the water, making sure that some or all of the phosphorus is above the surface. Ignite the phosphorus, and immediately cover the protruding part of the burning match stick with a vial and plunge the mouth of the vial below the surface of the water. Hold in this position until the phosphorus has burned out. Remove the matchstick from the vial while it is still under water, and, using the thumb as a stopper, remove the vial from the water. Shake it well and measure the length of the water column remaining and measure also the total length of the vial. Pour the water out and quickly thrust a burning splint into the remaining gas. Now compute the percentage of nitrogen and of oxygen in the air.

\[
\frac{\text{Length of water column}}{\text{Total length of vial}} \times 100 = \text{percent } O_2
\]

\[
100 \text{ percent } - \text{percent } O_2 = \text{percent } N_2
\]

2: Other Components

a: Carbon dioxide: On a glass plate, expose to the air a drop of clear limewater. Add water to prevent evaporation and observe this drop at intervals during the remainder of the class.

b: Water vapor: Fill two test tubes with ice water and tap water respectively and let them stand in the test tube rack. Also, expose a pellet of NaOH and another of CaCl₂ on a glass plate and set them aside until the next laboratory period. Observe the outside surface of the tubes of water.
II: WEIGHT AND VOLUME OF AIR

a: Heat a 4-in. test tube for a couple of minutes to drive out the air. Cork it immediately and let it cool. Now, immerse the corked tube in water and remove the cork. Repeat the test with the tube that has not been heated.

b: Weigh a test tube stoppered with a solid rubber stopper as carefully as you can. Remove the stopper and heat the tube for 2 minutes. Restopper it immediately and allow it to cool. Weigh it again.

III: REPORT FORM

Length of column of air (total length of collecting vial) ...................
Length of column of water ...........................................................
Percent of oxygen .................................................................
100 — Percent of oxygen, or percent of nitrogen .............................

Weight of tube of air ............................................................
Volume of tube ......................................................................
Weight of 1 ml of air ..................................................................
Density of air ..........................................................................
EXPERIMENT 41
FERROUS - FERRIC REACTIONS

I: EXPERIMENT

a: Place about 1 g of iron filings in a test tube or small flask. Add 5 ml H₂O and 2 ml conc. HCl. Mix and warm gently and let stand for 5 minutes. Dilute with 50 ml water and stopper the flask.

b: Take 2 ml of the stock solution prepared and put it in a small test tube. Add 3 drops of potassium ferricyanide. Repeat the process with three other tubes, adding, in turn, potassium ferrocyanide, potassium thiocyanate, and an excess of NH₄OH.

c: Take FeCl₃ from the stock shelf of reagents and repeat the whole of step (b) on this solution. Record the results of the Fe³⁺ and Fe²⁺ tests in parallel columns.

d: To 5 ml of the stock solution of FeCl₂ prepared above, add 1 ml conc. HNO₃ and 3 ml conc. HCl. Boil for about 3 minutes and repeat the tests of (b).

e: To 5 ml of FeCl₃ solution, add a gram of iron filings or scraps of iron and 2 ml conc. HCl. Boil until the color of the Fe³⁺ has disappeared and report the tests outlined above on this solution.

II: REPORT

Write reactions for all experiments done, making sure you have balanced equations, and label each oxidant and reductant.

REPORT EXPERIMENT HERE

<table>
<thead>
<tr>
<th>Fe⁺²</th>
<th>Fe⁺³</th>
</tr>
</thead>
</table>
1: PREPARATION OF SODIUM HYDROXIDE

Place a small piece of sodium, the size of a split pea, into a crucible one third full of water. STAND BACK WHILE THE REACTION IS TAKING PLACE. After all activity has ceased, remove a drop of the solution with a glass rod and test it with litmus paper. Add a drop of phenolphthalein or universal indicator to the solution in the crucible.

II: PROPERTIES AND TESTS FOR ALKALI METAL COMPOUNDS

In a small test tube place a pin head size crystal of NaCl. In other test tubes, do the same with NaNO₃, Na₂CO₃, NaHCO₃, NaC₂H₃O₂, KCl, KNO₃, KClO₃. Add 2 drops of water to each. Note the solubility of each and its reaction to litmus. (Recall the experiment on hydrolysis)

Flame tests for the alkali metal compounds are made as follows: Clean a piece of platinum or nichrome or chromel wire by dipping it in HCl and heating it in the flame until no color results. Dip the wire into the solution of the salt to be tested and hold the wire in the flame, noting the color it imparts to the flame. After cleaning the wire as previously directed, proceed to test the next solution in order and continue until all have been tested. This test is sensitive enough to detect 1/3,000,000 of a milligram of sodium.

For detecting the coloration of the flame produced by salts of potassium, hold two thicknesses of cobalt glass before your eyes as you look at the flame. This cuts out the sodium interference.

Obtain an unknown salt from the instructor and determine whether it is a salt of sodium or of potassium.
For any of the following experiments on metals that the teacher may choose to have the student do, it is suggested that the general outline report form for elements be followed. These reports are not done out in detail here as it is doubtful if many teachers will be interested in the section on metals. However, another suggestion for the reporting of metals follows:

<table>
<thead>
<tr>
<th>METAL MAY BE NAMED HERE</th>
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<tbody>
<tr>
<td>COLOR</td>
</tr>
<tr>
<td>MALLEABILITY</td>
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<td>DENSITY</td>
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<td>SOLUBILITY IN HCL</td>
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<td>EASE OF OXIDATION</td>
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<td>DUCTILITY</td>
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<td>METALLURGY</td>
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<tr>
<th>IMPORTANT ALLOYS</th>
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<th>IMPORTANT ORES</th>
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<th>STATES OR COUNTRIES PRODUCING</th>
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EXPERIMENT 43
THE ALKALINE EARTH METALS AND THEIR COMPOUNDS

I: FLAME TESTS

As in the previous experiment, make flame tests from solutions of CaCl₂, SrCl₂, BaCl₂. Obtain an unknown from the instructor and determine which of the cations is present.

II: PLASTER OF PARIS

Heat a large crystal of gypsum, or a spatulaful of gypsum powder slowly in a crucible and then set it aside to cool. When cool, transfer the material to a test tube and grind it well with a flattened glass rod. Place this powder on a glass plate and add a few drops of water until a paste is formed. Cover an oiled coin with this plaster of Paris paste and allow it to stand for about 20 minutes and then remove the coin.

III: HARD WATER

The presence of ions of Ca⁺² and Mg⁺² causes hardness in H₂O. If these ions can be removed by boiling, then the water is only temporarily hard, but if they are resistant to boiling, we have permanently hard water.

Place about 2 ml Ca(OH)₂ in a test tube and blow into it through a straw until a precipitate forms. Keep breathing until the precipitate re-dissolves. Label this solution “A”. This is temporary hard water because CaCO₃ dissolves in excess carbonic acid to form calcium bicarbonate.

Saturate 2 ml H₂O with CaSO₄, centrifuge, and pipet off the supernatant liquid into another tube and label it “B”. This is permanent hard H₂O.

Pour half of solution “A” and half of “B” into one test tube and label this mixture “C”.

Now, boil the original solutions “A” and “B”. Add about 5 drops of soap solution to each and shake. Add 5 drops of Na₂CO₃ to “B” which contains the permanent hard water, and shake.

Now take tube “C”, containing the mixture of hard waters and add 5 drops of dil. NaOH and 5 drops of soap solution. If suds do not form upon shaking, add more soap until they do. CaHCO₃ reacts with NaOH to produce CaCO₃, Na₂CO₃ and H₂O. The CaCO₃ precipitates and can be removed by centrifuging. The Na₂CO₃ reacts with permanent hard water (containing CaSO₄) to form CaCO₃ which can be centrifuged, and Na₂SO₄ which is a soluble salt having no hardening effect on water.
EXPERIMENT 44
MAGNESIUM AND ZINC

I: MAGNESIUM

1: Compounds of magnesium

Add a drop of 10 percent NaOH to a drop of MgSO₄ solution. Test with litmus paper.

Burn a small piece of magnesium ribbon, collecting the white ash on a glass. To this ash add a drop of water and test the resulting solution with litmus paper or phenolphthalein.

2: Test for the magnesium ion

Place a solution of MgSO₄ in a small test tube and make it slightly alkaline with dilute NH₄OH. Add a drop of Na₂HPO₄ and observe the formation of a white precipitate. In the absence of other metals, this is a reliable test for Mg⁺².

II: ZINC

1: Compounds of zinc

Dissolve a bit of granulated zinc in dilute HCl on a glass plate. Evaporate and observe the residue.

Dissolve this residue in a drop of H₂O and add dilute NH₄OH until a white precipitate forms. Add a drop of conc. NH₄OH and observe whether or not the precipitate is soluble.

2: Test for the zinc ion

Dissolve a bit of zinc chloride in some water, or use 2 ml of a zinc chloride solution in a test tube and make it alkaline with NH₄OH. Pass a few bubbles of H₂S and note the characteristically colored precipitate of ZnS that forms.
EXPERIMENT 45

ALUMINUM AND ITS COMPOUNDS

I: COMPOUNDS

In a small test tube, dissolve a small piece of Al foil in dil. HCl. Evaporate to dryness and dissolve the residue in H₂O. Make this solution alkaline with NH₄OH.

1: Amphoterism: To illustrate this property of aluminum hydroxide, take a small crystal of AlCl₃ in a test tube and by adding H₂O make some Al(OH)₃. Boil off any excess liquid and divide the residue into two parts. To one, add 10 percent NaOH until the precipitate dissolves, and to the other add dil. HCl until the precipitate in it dissolves.

2: Test for aluminum ion: Place a crystal of AlCl₃ in a test tube and dissolve it in H₂O. Make it alkaline with NH₄OH and observe the formation of a precipitate. This is the test for Al⁺³, but it can be made more sensitive by adding a dye such as alizarin or aluminon, before the precipitation of the hydroxide.

II: USES

1: Mordant: The gelatinous character of Al(OH)₃ makes it stick to fibers of cloth immersed in it. If a dye is added to the aluminum hydroxide, it also adheres to the fabric and is rendered thus “fast” to washing.

This may be illustrated with a piece of cotton cloth or yarn, and wool. Boil one piece of each in a small crucible containing 2 ml Al(C₂H₃O₂)₃ for a minute. Soak the cloth thoroughly and then boil it in 2 ml of alizarin for a minute. Wash each piece separately and allow to dry, then compare the colors.

2: Cleaning: When baking soda is boiled in aluminum pans, the pans are darkened, whereas if acid-containing foods are cooked in similar pots, the luster is increased. The mechanism of this reaction is not completely understood, but it is thought that the reaction is due to impurities, probably Fe, in the Al. These changes may be observed by heating a strip of Al alternately in 2 ml of water containing Na₂CO₃ and HC₂H₅O₂.

3: Goldschmidt Process: This method is used to reduce metals from their ores by the use of aluminum which combines with the oxide of the ore and frees the metal. Great amounts of heat are given off by the reaction.

On a piece of paper or glass, mix thoroughly about 0.1 g of Fe₂O₃ and the same amount of powdered aluminum. To this add carefully, being sure to mix, but not grind, 0.1 g of KClO₃ and 0.1 g Mg powder. Heat the mixture on an asbestos pad and insert the end of a piece of Mg ribbon into the heap. With the burner, ignite the free end of the Mg ribbon and step back until the reaction is complete. The Mg will carry the flame to the pile and the KClO₃ and Mg powder serve as fuse to ignite the whole mixture. Once the reaction has subsided, look for pieces of iron on the pad.
I: FLAME TESTS

For specific directions for these tests, refer to the experiment on alkali metals and alkaline earth metals.

II: BORAX BEAD TESTS

These tests are practically obsolete, but are included for those who may still wish to use them.

Practice making borax beads by filling a small loop in a platinum wire with borax. Heat this in the flame, and dip again into the borax. Once the bead is the size of the loop and clear, it is ready for use. Touch the warm bead to the substance to be tested and then hold in the oxidizing and in the reducing flame. Also, compare the color of the bead when hot and when cold.

<table>
<thead>
<tr>
<th>METAL ION</th>
<th>OXIDIZING FLAME</th>
<th>REDUCING FLAME</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>hot</td>
<td>cold</td>
</tr>
<tr>
<td></td>
<td>hot</td>
<td>cold</td>
</tr>
</tbody>
</table>

COPPER

IRON

COBALT

III: IDENTIFICATION OF SILVER AND LEAD

Mix 5 drops of Pb(NO₃)₂ and 2 drops of AgNO₃ in a test tube; add dil. HCl and centrifuge. Discard the supernatant liquid and add 5 drops of H₂O to the residue and stir. Warm the mixture and centrifuge again. Remove the supernatant liquid to another test tube and add to it 2 drops of K₂CrO₄ solution. If lead is present there will be a yellow precipitate of PbCrO₄. Again, treat the precipitate remaining in the centrifuge tube with hot water and again remove the supernatant liquid and add to it K₂CrO₄; continue this until the Pb⁺² test is no longer positive. Now, add conc. NH₄OH to the residue until it dissolves. This reaction is characteristic of AgCl. Acidify carefully with dil. HNO₃ and AgCl should again precipitate.

Report this experiment in the form of a flow chart of the reactions involved.
EXPERIMENT 47
THE IDENTIFICATION OF ARSENIC - ANTIMONY - BISMUTH —
GROUP 2B

I: EXPERIMENT
In each of three centrifuge tubes, place a spatulaful of powdered arsenic, antimony and bismuth. To each add a few drops of dil. HCl and centrifuge. Discard the undissolved metal and pass a stream of H₂S into the clear solutions. Centrifuge and note the color of the precipitates.

Using a known solution of any of these metals, pass H₂S through 2 ml of it and again note the color.

Obtain from the instructor either a powdered metal or a tube of unknown solution and proceed to analyze it. If there is any trouble in getting a sulfide precipitate, add a few drops of NH₄OH to the solution and add more H₂S.

II: REPORT
Report this analysis in the form of a flow chart of all the reactions involved.
EXPERIMENT 48

CHROMATOGRAPHIC ANALYSIS OF GROUP 1

i: EXPERIMENT

In separate crucibles, or in an evaporating dish, pour an 80-20 mixture of N-butyl alcohol and water until the vessel is about three quarters full. Cut strips of filter paper 1 x 10 cm. and fold them so that there is a sharp crease about 3 cm. from one end. Just below this crease, on the long side, place a streak of silver nitrate by drawing a stirring rod dipped in silver nitrate across the paper. On another strip, streak Pb(NO₃)₂ and on a third, HgNO₃. Place the short ends of these strips in the solvent in the evaporating dish and let the long side hang down. Allow the strips to remain in this position until the solvent has crept down about three quarters of the length of the paper, carrying ions of the salt with it. Remove the strips and brush or spray them with dil. K₂Cr₂O₇ solution. The migration of the metallic ions should be clearly visible by the color changes apparent on the strip. Repeat with one single strip streaked with all three metal ions. Obtain an unknown from your instructor and test it.

![strip of filter paper](image)

Fig. 27

ANALYSIS OF GROUP 2B

ii: EXPERIMENT

Repeat the operations above using salts of As, Sb, and Bi. As solvent you may try that prescribed above or you may have better success with the following mixture: 40 ml CH₃OH, 10 ml N-butyl alc., 10 ml H₂O, 10 ml isopropyl alc. This stock solution should supply several of you with enough solvent for the experiment. When the strips are ready for developing, brush or spray them with ammonium polysulfide solution, H₂S solution, or expose them to a stream of H₂S gas.

Again, repeat, using one strip streaked with all three metallic ions, and obtain an unknown from your instructor and analyze it. Check your analysis by repeating the previous experiment on the analysis of Group 2B salts.
I: DISTILLATION OF WOOD

Using a type "C" delivery tube, set up the apparatus as shown in the diagram.

![Diagram of distillation setup]

Place sawdust in the generator until it comes up about a fourth of the way in the tube, then slant the tube into position on the ring stand so that there is an air space above the sawdust. Heat the tube gradually until there is a hot flame directly beneath the sawdust and the tube is red hot. Continue heating thus until there is a bit of distillate in the collecting tube. Test this substance with litmus paper and try its solubility in water. Examine the material remaining in the generator.

II: DISTILLATION OF COAL

This is carried on in the same manner as the distillation of wood, but in place of sawdust, powdered coal is substituted.
EXPERIMENT 50
HYDROCARBONS: METHANE - ACETYLENE - GASOLINE - KEROSENE: PREPARATION AND PROPERTIES

I: PREPARATION OF METHANE

Using a type "A" set-up for the collection of a gas, place a mixture of 0.1 g fused NaC₂H₃O₂ and 1.5 g soda lime, well mixed together by grinding, into the generator. Heat the generator and collect two vials of methane by the displacement of water. In a third vial, only half filled with water, displace this much water with the methane thus giving a vial of air and methane mixed.

Properties

Test the combustibility of the gas in the first vial and also test the products of combustion by means of limewater and cobalt chloride paper. If you do not get a positive limewater test, try it over again with another vial of gas.

Try igniting the gas that was contaminated with air. This should produce a mild explosion on ignition.

II: PREPARATION OF ACETYLENE

Place a small lump of CaC₂ and H₂O in an evaporating dish. Into this invert a collecting vial of water. The CaC₂ and H₂O react to give acetylene which will displace the water in the collecting vial. Put a little water in another vial and displace it by acetylene. Test the combustibility of both the pure and the impure acetylene.

III: GASOLINE AND KEROSENE

Both of these substances ignite easily and have low kindling points. Your teacher may wish you to light a drop of either on a glass slide. To play with larger quantities is unsafe.
EXPERIMENT 51
DERIVATIVES OF METHANE — PROPERTIES

I: PROPERTIES

a: Compare the solubility in water of $\text{CCl}_4$, $\text{CS}_2$, and $\text{CHCl}_3$. Compare the densities of these substances with each other and with water.

b: Test the combustibility of each by applying a burning splint to a drop of the substance on a glass plate.

c: Retest the solubilities using ethyl alcohol as the solvent instead of water.

d: Tabulate the results of your experiment in the following chart:

<table>
<thead>
<tr>
<th>SUBSTANCE</th>
<th>FORMULA</th>
<th>VOLATILITY</th>
<th>FLAMMABILITY</th>
<th>DENSITY COMPARED WITH $\text{H}_2\text{O}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>CARBON DISULFIDE</td>
<td></td>
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</tr>
<tr>
<td>CARBON TETRACHLORIDE</td>
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</tr>
<tr>
<td>CHLOROFORM</td>
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</table>

<table>
<thead>
<tr>
<th>SUBSTANCE</th>
<th>SOLUBILITY IN $\text{H}_2\text{O}$</th>
<th>SOLUBILITY IN C$_2$H$_5$OH</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{CS}_2$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{CCl}_4$</td>
<td></td>
<td></td>
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<tr>
<td>$\text{CHCl}_3$</td>
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</table>

II: PREPARATION OF IODOFORM

Place 3 drops of alcohol in a centrifuge tube and add 1 drop of conc. NaOH. Heat gently while adding dilute iodine solution dropwise until a yellow precipitate appears. The same test may be applied to acetone but will not work with methyl alcohol, but it would be well for you to try both the acetone and the methyl alcohol in place of the ethyl alcohol as given in the above directions.
EXPERIMENT 52
CARBOHYDRATES — SUGAR

I: SUCROSE

Heat a few crystals of sugar in a small test tube and note the change that occurs in the sugar and the deposit on the walls of the tube.

Invert sugar may be formed by the following process: In a large test tube, add 0.1 g sugar to 2.5 ml H₂O. Add a drop of dil. HCl and boil for a minute. Neutralize with NaOH. Test with Fehling’s solution or Benedict’s solution. If Fehling’s solution is used, add 1 ml Fehling solution A to 0.5 ml Fehling solution B and add 1 ml of this mixture to the solution to be tested. Warm gently until a precipitate forms. If Benedict’s solution is used, add 1 ml of it and heat the mixture gently until the precipitate forms.

II: GLUCOSE

Determine the physical properties of glucose or dextrose. Compare its solubility with that of sucrose in water by adding amounts of each to a fixed quantity of H₂O in separate test tubes.

Test glucose with Fehling’s solution.

In a small clean test tube, place 0.3 ml AgNO₃ solution and add dil. NH₄OH until the precipitate which forms disappears. Add 5 drops of glucose solution. Warm the tube gently until a silver mirror is formed.

Tiny bits of candy may be dissolved and tested for glucose in the same way. Other sweet products may be used, such as molasses or honey.
I: PREPARATION OF STARCH

Grind a small piece of potato to a fine pulp and transfer it to a centrifuge tube. After centrifuging, discard the supernatant liquid and test the residue in the following ways:

a: *Iodine test*: Place a bit of the starch on a watch glass, moisten it with a drop of water, and add a drop of iodine solution. The formation of a blue color is a positive test for the presence of starch. You may test wheat, corn, or products containing these in the same way.

b: *Hydrolysis of starch*: In digestion, starch is converted into sugar by the presence of acids in the digestive tract. You may hydrolyze starch in the laboratory by taking a small quantity of your starch paste in a small test tube. Add to it a drop of conc. HCl and heat until you can no longer get a positive test with iodine. (To determine when this point is reached, remove a speck of the starch mixture with a stirring rod and test it. Repeat until the test is negative.) When you are sure that all the starch is removed, make a Fehling’s sugar test on the converted sugar solution.

II: CELLULOSE

a: Cotton, wood, and other vegetable products are rich in cellulose. Cotton is almost 100 percent cellulose. Test a bit of it with iodine solution. Put a small chip of wood in a test tube and add to it conc. H$_2$SO$_4$ and observe the results.

b: *Hydrolysis of cellulose*: In this hydrolysis, we are illustrating a step further back; we first hydrolyze cellulose to starch and then carry the starch on to sugar, all by the aid of HCl. Place a spatulaful of sawdust or cotton the size of a pea, into a crucible containing 5 ml H$_2$O and 2 drops conc. HCl. Boil for 10 minutes, replacing the water lost by evaporation. Test a bit adhering to the stirring rod with iodine solution. Once the test for starch is obtained, 10 minutes more of boiling should hydrolyze the starch to sugar. Neutralize the excess acid and test for glucose by adding 5 drops of Fehling’s or Benedict’s solution. Heat until the color changes from blue to orange.
EXPERIMENT 54
ESTERS AND SOAPS

I: PREPARATION OF AN ESTER

Place about a gram of salicylic acid in a small beaker and add 15 ml CH₃OH and cautiously pour 5 ml conc. H₂SO₄ down the side of the beaker. Heat the mixture under a hood for about 15 minutes and then carefully smell it. Add H₂O and see if you can detect oily droplets forming. These are the ester.

II: PREPARATION OF SOAPS

Place a small lump of some hydrogenated fat, 1 g of KOH and 15 ml CH₃OH in a small beaker and heat gently in the hood for 15 minutes. Allow to cool and then introduce a few drops of the solution into a test tube. Add a little water and shake well. In another tube place a few drops of the soap solution and add CaCl₂ until no further change occurs. Place a few drops of the soap solution in a test tube and mix with H₂O. Add HCl until no further change occurs.

III: Alternate methods of saponification may be discovered by consulting other text books. If your teacher would like you to try one of these in preference to the above, you will find it good practice.
EXPERIMENT 55
RADIOISOTOPES

I: DISCUSSION

The solutions of KI in this experiment are made in the identical manner, but one tube has a bit of radioactive iodine in the form of radioactive KI. Throughout the chemical history of this substance, to be traced in this experiment, it will be found that I\textsubscript{131} behaves in identically the same manner as the normal isotope I\textsubscript{127}. This proves that isotopes are the same chemically but differ only in the physical property of atomic weight. Their extranuclear electronic configurations are identical.

II: EXPERIMENT

Place 5 ml of 2N solutions of KI in each of 2 small test tubes. Add some radioactive KI containing about 4 microcuries of I\textsubscript{131} to one of the tubes. Add Pb(NO\textsubscript{3})\textsubscript{2} in excess to both tubes to insure complete precipitation of PbI. Centrifuge and wash the precipitates thoroughly with at least two wash waters. Test the precipitate with a Geiger Counter. Test the filtrates with the counter also.

Dissolve the two precipitates in saturated Na\textsubscript{2}CO\textsubscript{3} which will result in the formation of white lead carbonate from yellow lead iodide. Centrifuge this solution. Remove the centrifugate and test it with a Geiger Counter. Wash the precipitate twice and test it with the Counter.
EXPERIMENT 56

RADIOACTIVE TRACERS

I: DISCUSSION

Using a solution containing radioactive I$^{131}$ in a 2N KI solution, the migrations of the iodine atoms can be followed by the use of a Geiger Counter on these tagged atoms. The reaction is slow and therefore may lose some of its effect on the student mind, but the point of diffusion can be well taught with this tool.

II: EXPERIMENT

In a large cylinder, place 200 ml of 2N KI. Introduce 10 ml 2N KI containing 10 microcuries of I$^{131}$, into the bottom of the cylinder by means of a long pipet. Record the background reading of the Geiger Counter. Immediately take readings at 2 inch intervals up the cylinder. Repeat these readings at the same levels on the 1st, 2nd, 3rd, 4th, and 5th days following. On the last day, stir the contents of the cylinder and take a reading after that.

III: REPORT

Record all readings on the following chart:

<table>
<thead>
<tr>
<th>LEVEL</th>
<th>INITIAL</th>
<th>1 DAY</th>
<th>2 DAYS</th>
<th>3 DAYS</th>
<th>4 DAYS</th>
<th>5 DAYS</th>
<th>STIRRED SOLUTION</th>
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I: DISCUSSION

Milk consists of fats, protein and other nitrogen compounds, carbohydrates, vitamins and inorganic compounds. By careful laboratory work you can identify the following:

II: EXPERIMENT

1: FAT: To 2 ml of whole milk, add 1 ml of conc. H₂SO₄ and mix thoroughly with a stirring rod. Centrifuge for one minute. Add 1 ml of hot water and centrifuge again. The formation of a layer of fat on the top of the liquid in the tube is evidence of the fat present.

2: CASEIN: To 2 ml of whole milk, add 1 ml H₂O, and add slowly 1 ml 10 percent solution of acetic acid. Let the mixture stand 3 minutes and then either filter through glass wool or centrifuge, and retain both the filtrate and the precipitate. Test the precipitate for protein by adding 10 drops of conc. HNO₃. A yellow color is a positive test for a protein.

3: LACTALBUMEN: Heat one-half the filtrate from the casein experiment until it boils. This coagulates the lactalbumen. Filter through glass wool or centrifuge. Test the residue for protein as you did before. Lactalbumen is protein that is not coagulated with acids.

4: LACTOSE: Test the filtrate from the lactalbumen separation for a reducing sugar by adding 2 ml of Benedict’s solution and heating gently. A yellow-green precipitate indicates the presence of lactose.

5: PHOSPHATE: To the filtrate obtained in the casein separation, add 1 ml of ammonium molybdate and heat gently. The formation of a yellow flocculent precipitate is indicative of the presence of the phosphate ion.

REPORT YOUR EXPERIMENT HERE.
EXPERIMENT 58

COMPOSITION OF BAKING POWDERS

I: DISCUSSION

Baking powders are of several types, all used for the purpose of leavening, but containing different substances in addition to carbonates. Sodium bicarbonate, starch, cream of tartar, potassium alum and monocalcium or disodium phosphate are the most commonly found substances. Obtain samples of two or more types of baking powder and proceed to identify the compounds contained therein.

II: EXPERIMENT

1: Sodium bicarbonate: Fill the curved part of a small test tube with baking powder. Add 2 ml H₂O and shake the tube vigorously. Perform a flame test on this solution by dipping the end of a platinum wire into the solution and then holding the wire in the flame of a burner. A yellow flame indicates the presence of sodium, and will assure you that sodium is there compounded, in all probability, with the bicarbonate radical. By looking at the flame again, through two thicknesses of cobalt glass, you can ascertain whether or not potassium is also present. A violet shade, seen through the cobalt glass, is indicative of potassium ions. The potassium may be part of a cream of tartar or an alum salt.

   Fill the rounded part of a test tube with baking powder again, and add 2 ml conc. HCl. Suspend a drop of limewater from a glass rod and hold this in the escaping gas. A milky appearance developing in the limewater is a positive test for the presence of the carbonate or bicarbonate ion.

2: Starch: Fill the rounded part of a test tube with baking powder and add 2 ml H₂O. Gently boil the mixture for 2 minutes. Add a drop of tincture of iodine and if a dark blue color is formed, starch is present.

3: Cream of tartar: If potassium has been found present, it is possible that the tartrate radical is also in the powder. Fill the curved part of a test tube with baking powder and add 3 drops of conc. H₂SO₄, and heat the mixture. Formation of a charred mass which smells like burned sugar is evidence for the presence of tartrate.

4: Alum: Alums are aluminum sulfates which may be present in the baking powder as the potassium or ammonium salt. To 10 drops of centrifugate obtained by centrifuging the solution remaining after the CO₂ has been driven off in part (1), add 2 drops of NH₄OH. The formation of a white, gelatinous precipitate indicates the presence of Al³⁺.
Dissolve a bit of the baking powder in water and centrifuge. Take the clear solution and add 3 drops of BaCl₂. A white precipitate of BaSO₄ indicates the presence of the SO₄⁻². However, interfering ions might produce this white precipitate, so, add HCl to it, and if it remains insoluble, it is a sulfate.

To a bit of the baking powder, add 2 ml 20 percent NaOH. Heat the tube and hold a piece of moist red litmus at the mouth. If ammonia is present in the escaping vapors it will turn the litmus blue.

5: Monocalcium phosphate, or disodium phosphate: Take a bit of baking powder and add 2 ml of water and shake well. Centrifuge if necessary. Add glacial acetic acid until the solution is acid to litmus and then add 2 ml of ammonium oxalate solution. A white precipitate, which is often slow in appearing, is indicative of calcium.

Take a bit of baking powder and add 2 ml dil. HNO₃. Centrifuge and to the clear liquid, add 2 ml ammonium molybdate. Warm the tube and allow it to stand. The formation of a yellow precipitate is a positive test for a phosphate.

REPORT THE RESULTS OF YOUR TESTS HERE
AND WRITE REACTIONS EQUATIONS
EXPERIMENT 59
ACIDITY OF COMMON FOODS

I: DISCUSSION
The quantitative neutralization of a substance by a base of known concentration will indicate the acidity of that substance. Using 0.1 N NaOH, determine the acidity of the following foods:

II: EXPERIMENT
1: Ice cream: Dilute 2 ml of soft vanilla ice cream with 2 ml H₂O. Mix thoroughly in an Erlenmeyer flask and add 5 drops of phenolphthalein. Titrate with 0.1 N NaOH until the first permanent pink color results. For calculations, see Experiment XXI on vinegar.

2: Carbonated beverage: Dilute 2 ml of any carbonated beverage with water and mix well in an Erlenmeyer flask. Add 5 drops of phenolphthalein and titrate as before.

3: Orange juice: Pipet 2 ml of clear orange juice into an Erlenmeyer flask, dilute with water, and titrate to the phenolphthalein endpoint with 0.1 N NaOH. (If the orange juice is not clear, centrifuge or filter it and use the clear liquid resulting.)

4: Other foods may be treated in a similar manner to determine their acidity. They must first be liquified, and then titrated. It is suggested that you work with tomato and other fruit juices, and milk.

REPORT EXPERIMENT HERE
ANALYSIS OF TOOTH POWDER

I: DISCUSSION

A knowledge of methods of chemical analysis will often help you to discern false claims in advertising. In the general field of cosmetics, the Pure Food and Drug Laws protect the consumer to a large extent, but even with legal protection, some adulterants or even harmful ingredients may find their way into these preparations. It is suggested that you find two or three brands of dentifrices commercially available, and compare the results of tests you make on them.

II: EXPERIMENT

1: Test for CO$_3^{2-}$: Place a spatulafull of tooth powder in a test tube and add 2 ml of dil. HCl. Connect a type “C” delivery tube and dip the end into a tube containing 2 ml Ca(OH)$_2$. If CO$_2$ gas is escaping, the lime water will assume a milky appearance.

2: Test for PO$_4^{3-}$: Centrifuge the acid-powder solution used above and pipet 5 drops of the clear solution into a test tube. Add NH$_4$OH dropwise until the solution is neutral to litmus. Add 1 drop of HNO$_3$ and a milliliter of ammonium molybdate. Warm the solution gently and wait a few moments for the formation of a yellow precipitate of ammonium phosphomolybdate. If phosphate is present, you have a tooth powder containing a harmful abrasive.

3: Acidity: Place a spatulafull of tooth powder in a test tube and add 2 ml H$_2$O. Add 1 drop of universal indicator, shake and centrifuge. Determine the pH by comparing the color with the chart showing pH range. Repeat, using alcohol as a solvent in place of water. The possibility of hydrolysis is thus prevented, and you have a measure of the actual free acids present in the tooth powder.

4: Calcium: To a milliliter of centrifugate from step 1, made neutral with NH$_4$OH, add 1 ml of ammonium oxalate. A precipitate indicates the presence of calcium.

5: Starch: Place a few grains of tooth powder on a watch glass, and moisten with a drop of tincture of iodine. A dark blue color indicates the presence of starch.

6: Soap: Place a few grains of tooth powder in a test tube, add 2 ml H$_2$O and shake well. If suds form, soap is present.

7: Abrasives: Remove the residue from part 1 and place it on a clean glass plate the surface of which is perfectly smooth. Cover this with another glass plate and rub the two together for a full minute. Remove and wash the plates and examine them for scratches.
QUANTITATIVE DETERMINATION OF VITAMIN C

I: DISCUSSION

Vitamin C, or ascorbic acid, is a common ingredient of all citrus fruits and an important vitamin in maintaining good health. You might like to find which foods contain the most vitamin C. Perhaps you do not know whether there is more of this vitamin in a glass of fresh orange juice, canned orange juice or frozen orange juice. If you obtain samples of all of these, you can find out for yourself which source is richest.

II: EXPERIMENT

Pipet 2 ml of clear orange juice into a 125 ml Erlenmeyer flask and add 10 ml H₂O and 1 ml of 1 percent starch solution. If the orange juice you have is not clear, centrifuge or filter it and use the clear liquid that is thus prepared.

Titrate the mixture rapidly with an accurately standardized 0.01 N solution of iodine containing 16 g of KI per liter of solution. Each ml of iodine is equivalent to 0.85 mg of ascorbic acid.

REPORT EXPERIMENT HERE
Classify each of the following by checking the proper column:

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CHEMICAL NOMENCLATURE

Place the chemical name next to the formula:

KBr  HNO3
FeCl3 PbCrO4
HNO3 PbS
KMnO4 HCl
KNO3 LiNO2
H3BO3 PbO
MgSO4 SnCl2
NaOH CuO
BaSO4 CO2
P2O5 MnO2
CaF2 Na2B4O7
H2SO4 NaClO3
CCl4 FeCl3
PbSO4 H3PO4
CCl4 Na2SO4
PbO2 Fe2O3
KNO3 H2O
As2S NaOH
AgCl SiC
PbSO4 (NH4)2SO4
HgCl2 H2PO3
H2S MgSO3
H2S SnCl4
Na2O5 Fe2(CO3)3
Ag2O BaO
As2O3 BaO2
HgCl2 BaSO4
HClO3 AgClO3
ZnCO3 Sn (C2H5O2)4
ZnSO4 NaCN
K2SO4 KCNS
ZnO (NH4)2P2O7
ZnCl2 HClO2
CuSO4 N2O5
ZnCl2 N2O
Na2S03 NO
Na2S04 N0
**ATOMIC STRUCTURE**

Draw the Bohr-Rutherford diagrams for the following:

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<th>fluorine atom</th>
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<tbody>
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<td>chlorine atom</td>
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</tr>
<tr>
<td>hydrogen atom</td>
<td>aluminum</td>
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<td>carbon</td>
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110
Fill in the correct formulas and practice naming the compounds:

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Change each given unit to milligrams, millimeters or milliliters, centimeters, centigrams or centiliters, grams, meters or liters and kilograms, kilometers or kiloliters, as the case may be and enter the results in the proper columns.
WRITING FORMULAS FOR COMPOUNDS

Place the appropriate formula next to the compound:

antimony chloride
potassium hydroxide
sodium nitrate
silver sulfide
sulfuric acid
potassium chlorate
chlorous acid
tungstic acid
francium fluoride
sodium iodide
scandium orthosilicate
magnesium sulfite
manganese dioxide
calcium astatide
hydrogen iodide
sodium bicarbonate
cuprous chloride
ferric nitrate
arsenic sulfide
stannous iodide
calcium oxide
potassium dichromate
nitric acid
hydrosulfuric acid
cobaltous selenide
calcium acetate
ammonium hydroxide
nitrous oxide
cobaltic nitrate
sodium hydroxide
ferric nitrate
bismuth oxide
magnesium bisulfate
sodium acid sulfate
potassium thiocyanate
ferric ferrocyanide
aluminum hydroxide
Using a chart of atomic weights, calculate the molecular weights of the following:

- CuO
- SO₂
- H₂
- KNO₃
- NH₄OH
- (NH₄)₃PO₄
- Al₂(O₃)₃
- CuSO₄·5H₂O
- Na₂S₂O₃·5H₂O
- KClO₃
- CuCO₃
- Fe₂(CO₃)₃
- H₃S
- (NH₄)₂SO₄
- CH₃COOH
- Pb(C₂H₃O₂)₃
- H₂SO₄
- NaOH
- NH₄OH
- Ca(OH)₂
- H₃PO₄
- ZnO
- CaCO₃
- N₂O
- Al(OH)₃
BALANCE THE FOLLOWING EQUATIONS:

1: \( C + O_2 \rightarrow \)
2: \( Fe + O_2 \rightarrow \)
3: \( P + O_2 \rightarrow \)
4: \( Fe^+ + O_2 \rightarrow \)
5: \( N_2 + H_2 \rightarrow \)
6: \( Al + NaOH \rightarrow \)
7: \( Fe + H_2O \rightarrow \)
8: \( CH_4 + O_2 \rightarrow \)
9: \( KClO_3 \rightarrow \)
10: \( Cu + H_2SO_4 \rightarrow \)
11: \( Na_2SO_4 + H_2SO_4 \rightarrow \)
12: \( CaCO_3 + HCl \rightarrow \)
13: \( AgNO_3 + HCl \rightarrow \)
14: \( BaCl_2 + Na_2SO_4 \rightarrow \)
15: \( K_2CrO_4 + PbCl_2 \rightarrow \)
16: \( As(NO_3)_3 + H_2S \rightarrow \)
17: \( ZnCO_3 + H_2S \rightarrow \)
18: \( PbSO_4 + (NH_4)_3S \rightarrow \)
19: \( CaO + C \rightarrow \)
20: \( H_2O \) electrolysis \( \rightarrow \)
21: \( Mg + HCl \rightarrow \)
22: \( Fe_3O_8 + Al \rightarrow \)
23: \( AgNO_3 + H_2S \rightarrow \)
24: \( Pb(NO_3)_2 + K_2CrO_4 \rightarrow \)
25: \( Al_2(SO_4)_3 + NH_4OH \rightarrow \)
26: \( CuO + H_2 \rightarrow \)
27: \( Fe_2O_3 + C \rightarrow \)
28: \( K_2Fe(CN)_6 + FeCl_2 \rightarrow \)
29: \( K + H_2O \rightarrow \)
30: \( (NH_4)_2SO_4 + KOH \rightarrow \)
31: \( NaNO_3 + H_2SO_4 \rightarrow \)
32: \( MnO_2 + HCl \rightarrow \)
33: \( Cu + Cl_2 \rightarrow \)
34: \( SbCl_5 + H_2S \rightarrow \)
35: \( Pb(NO_3)_2 + H_2S \rightarrow \)
36: \( CaC_2 + H_2O \rightarrow \)
37: \( Cr_2O_3 + Al \rightarrow \)
38: \( Cu(OH)_2 + H_2S \rightarrow \)
EQUATION BALANCING — continued

39: NaI + Br₂ →

40: Na₂CO₃ + AgNO₃ →

41: NaCl + AgNO₃ →

42: Sb + Cl₂ →

43: H₂SO₄ + Fe →

44: Al + HCl →

45: CaCO₃ + HCl →

46: HC₂H₃O₂ + Ca(OH)₂ →

47: H₂SO₄ + NaOH →

48: Na₂CO₃ + H₂SO₄ →

49: CaO + H₂O →

50: NaCl + HNO₃ →

51: Fe₃O₄ + CO →

52: K + HCl →

53: Ca + H₂O →

54: CaSO₄ + Na₂CO₃ →

55: MgBr₂ + Cl₂ →

56: NaHCO₃ + HCl →

57: KHSO₄ + H₂SO₄ →

58: Na₂C₂H₃O₂ + H₂SO₄ →

59: KCl + H₂SO₄ →

60: Cu + conc. HNO₃ →

61: Cu + conc. H₂SO₄ →

62: Fe₃O₄ + Al →

63: Cu + dil. HNO₃ →

64: Ca(OH)₂ + HNO₃ →

65: Al + O₂ →

66: CuSO₄ + H₂S →

67: Zn(NO₃)₂ + H₂S →

68: HgO →

69: CaCO₃ + H₂CO₃ →

70: Ca(HCO₃)₂ →

71: Na₂SO₃ + HCl →

72: NH₄HCO₃ + NaCl →

73: NH₃ + H₂O →

74: NH₃ + HCl →

75: SO₂ + H₂O →
1: Calculate the percentage composition of the following compounds:
   water
   mercuric oxide
   ferric oxide
   gypsum
   sulfuric acid
   phosphorus pentoxide
   zinc sulfate
   crystalline blue vitriol
   crystalline potassium sulfate
   borax

2: What percent oxygen is contained in crystalline K₂SO₄·MgSO₄·6H₂O?

3: What percent of KClO₃ is oxygen? How much O₂ will be formed from 100 g of it?

4: How much arsenic is available in 100 lbs. of realgar, containing 70 percent arsenic?

5: How much phosphorus can be obtained from a ton of calcium phosphate?

6: Which will give more copper, an ore containing 95 percent CuCO₃, or an ore containing 90 percent Cu₂S?

7: Determine the percentage composition of difluorodichloromethane (CF₂Cl₂).

8: A mine produces malachite (Cu₂(OH)₂CO₃) and azurite (Cu₃(OH)₂(CO₃)₂). Which ore is richer in copper and how much copper could be obtained from 2 tons of ore?

9: What percent of sulfuric acid is sulfur? What percent of sulfurous acid is sulfur?

10: Calculate the percent of CaO in CaCO₃. How much CaO (in pounds) can be obtained from 1 ton of limestone that is 97 percent CaCO₃?

11: A sample of impure sulfide ore contains 42.34 percent zinc. What is the percent of zinc sulfide in the sample?

12: A 5.82 g silver coin is dissolved in HNO₃. When NaCl is added to the solution, all the silver is precipitated as AgCl. The AgCl weighs 7.20 g. Determine the percent of silver in the coin.

13: How many tons of Ca₅(PO₄)₂ must be treated with carbon and sand in an electric furnace to make 1 ton of phosphorus?

14: How many tons of iron could be obtained from 2,000 tons of hematite, considered as pure Fe₂O₃?
If the density of air is 1.29 (STP), what is the density of each of the following gases in relation to air?

NITROUS OXIDE
NITRIC OXIDE
NITROGEN PENTOXIDE
OXYGEN
HYDROGEN
NITROGEN
ACETYLENE
AMMONIA
DIFLUORODICHLOROMETHANE
PHOSGENE (COCl₂)
ETHANE
CARBON DIOXIDE
CARBON MONOXIDE
SULFUR DIOXIDE
HYDROGEN CHLORIDE

If 200 ml SO₂ weigh 0.57 g, what is the density of the gas?

If an alloy contains 7 parts by weight of lead, (11.35 g/ml) and 25 parts tin, (7.3 g/ml) what is its density?

The weight of 22.4 liters of a gas is 17 g. What is its density?

A stone occupies 50 ml and its weight is 165 g. What is its density?

A substance weighs 448 g and displaces 40 ml H₂O. What is its density?

250 ml of SO₂ weigh 0.72 g. Calculate the weight of 1 liter, the molecular weight, the vapor density and the specific gravity.

If 0.74 g of CO₂ occupy 370 ml, find the weight of 1 liter, the vapor density, the molecular weight and the specific gravity.
DENSITY PROBLEMS

1: The molecular weight of alcohol vapor is 46. Calculate the vapor density, the specific gravity, and the weight of 1 liter of it.

2: The vapor density of HCl is 18.25. What is the molecular weight, the specific gravity, and the weight of 1 liter?

3: Ammonia gas had a vapor density of 8.5. What is the specific gravity, the molecular weight, and the weight of 167 ml?

4: 500 ml of phosphorus vapor weigh 2.79 g. What is the weight of 1 liter? Find the vapor density, the molecular weight and the specific gravity.

5: 100 ml of a gas weigh 0.228 g. Find the molecular weight.

6: The molecular weight of a gas is 71. What is the weight of 1 liter?

7: What is the weight of 1 liter of CO?

8: 750 ml of methane weigh 0.54 g. What is the molecular weight?

9: The specific gravity of CH₄ referred to air is ? To O₂? the vapor density?

10: Determine the weight of 50 ml of glycerine, sp. gr. 1.26.

11: Determine the volume, in gallons, of 400 lbs. of cottonseed oil which has a specific gravity of 0.926. One gallon of water weighs 8.34 lbs.

12: Calculate the weight in pounds of 3.00 cubic feet of glycerine, sp. gr. 1.26.

13: Battery acid has a specific gravity of 1.285 and contains 38 percent by weight of H₂SO₄. How many grams of pure H₂SO₄ are contained in a liter of battery acid?

14: Calculate the weight of pure HNO₃ per milliliter of the conc. acid which assays 69.8 percent by weight HNO₃ and has a sp. gr. of 1.42.

15: A piece of plastic foam 13x9.5x2.5 inches weighs 350 g; a cellulose sponge 7x12x25 cm. weighs 12 g. Foamglass is light as balsa, 10 lbs./ft³. It is ½ as heavy as cork. Calculate the specific gravity of each of the plastics and of cork.

16: A sample of conc. H₂SO₄ is 95.7 percent H₂SO₄ by weight and has a sp. gr. of 1.84. How many grams of pure H₂SO₄ are contained in 1,000 ml of acid? How many milliliters of acid contain 100 g of pure H₂SO₄?

17: Determine the weight of 20 ft³ of aluminum.

18: Find the volume of 40 kg. CCl₄.
Given the following information, determine the formula of the compound described.

1: M.W. 230; 62.5 percent arsenic; 34.8 percent oxygen.
2: M.W. 30; 80 percent carbon; 20 percent hydrogen.
3: M.W. 58; 82.75 percent carbon; 17.25 percent hydrogen.
4: No molecular weight given; contains 7.7 percent hydrogen and 92.3 percent oxygen.
5: An acid contains 40 percent carbon, 6.6 percent hydrogen and 53.4 percent oxygen.
6: 96.75 percent of the compound is iodine and 3 percent hydrogen.
7: An oxide of barium has 81 percent barium and 19 percent oxygen.
8: A compound contains 1 percent hydrogen, 11.99 percent carbon, 47.95 percent oxygen and 39.06 percent potassium.
9: 168 pounds of iron oxidize to 240 pounds of iron oxide. Find the formula of the oxide.
10: What is the formula of the compound having 1.832 g of copper and 0.462 g of oxygen?
11: An oxide of sulfur contains 40 percent sulfur and 60 percent oxygen.
12: An oxide of iron contains 70 percent iron and 30 percent oxygen.
13: 3.06 percent hydrogen, 31.63 percent phosphorus, 65.30 percent oxygen.
14: 10 g of pure tin produce 12.7 g of the oxide. Find the formula of the oxide.
15: From 39 g of an oxide of mercury, 1.5 g of oxygen are prepared. Find the formula of the oxide.
16: The simplest formula of a gas is CH₂. Its molecular weight is 28. What is its formula?
17: A hydrocarbon has the following composition: 92.31 percent carbon and 7.67 percent hydrogen. The molecular weight is 78. Find the formula.
18: Find the correct formula for the hydrate with the composition Ca 18.27 percent; chlorine 32.42 percent; water 49.31 percent; molecular weight 219.
VOLUME - VOLUME PROBLEMS

1: What is the volume of air needed to burn 1 liter of C₂H₅OH?
2: What volume of oxygen is needed to burn 175 ml of acetylene?
3: What volume of oxygen will be required to burn 40 liters of benzene?
4: What volume of CO₂ is formed when 15 liters of ethane are burned?
5: What volume of HCl is formed when 50 liters of Cl₂ combine with H₂?
6: 100 ml H₂ and 175 ml Cl₂ are exposed to light. What is the volume of the resulting gas?
7: 1 volume of CO and 1 volume O₂ are exploded. What is left? What volume of CO₂ is produced?
8: How much O₂ is needed for the complete combustion of 15 liters of H₂S?
9: 10 liters of H₂ and 30 liters of Cl₂ are exploded. What is left and how much?
10: 100 liters of N₂ are combined with H₂ and produce how much ammonia?
11: 100 ml H₂ + 100 ml O₂ react. Which gas is in excess and by how much?
12: 100 ml CO and 150 ml CH₄ and 1,000 ml O₂ react. What is formed and how much?
13: 14 liters of CO produce how much CO₂ upon oxidation?
14: How much Cl₂ can be made from 1 liter of HCl?
15: 150 liters of O₃ + Sulfur produces how much SO₂?
16: What volume of H₂ will combine with 12 liters of Cl₂ to form HCl? What volume of HCl will be formed?
17: A commercial method for the preparation of NO is by the oxidation of ammonia in the presence of a catalyst. How many liters each of O₃, NO and H₂O vapor are used or produced when 10 liters of ammonia are oxidized?
18: What volume of O₃ is required for the complete combustion of 1 mole of CS₂? What volumes of CO₂ and SO₂ are produced?
19: 50 liters of air will burn with how much methane?
20: 20 liters of benzene are burned in air. How much air is required for the combustion?
1: If 98 g of KClO₃ are heated, what weight of KCl results?
2: If dil. H₂SO₄ is added to 350 g of Na₂CO₃, what weight of CO₂ is evolved?
3: What weight of CO₂ is formed from burning 1 ton of coke?
4: 48 g of Mg react with HCl to yield a white solid that is recoverable by evaporation. What is it and how much of it is there?
5: 3.9 g of potassium react with water to produce how much KOH?
6: Zinc and 355 g of Cl₂ are warmed together. What weight of ZnCl₂ is formed?
7: Al₂O₃ is decomposed into its elements in a furnace. How much ore is needed to produce 275 lbs. of metal?
8: If hematite ore is 75 percent Fe₂O₃, how much must be mined to yield 5 tons of iron after reduction with coke?
9: What weight of aluminum is needed to produce 75 g of aluminum nitride by direct combination?
10: What weight of H₂SO₄ will react with 45 g of aluminum completely?
11: How much HCl will yield 5,000 g of Cl₂?
12: What weight of Fe will release 100 g of H₂ from H₂SO₄?
13: What weight of salt will produce 4,500 g of HCl?
14: 50 g KClO₃ and 50 g NaClO₃ are heated separately. Which will produce the greater amount of O₂?
15: How much MnO₂ is needed to get 250 g of Cl₂ from HCl?
16: How much FeS must react to produce 400 g of H₂S?
17: How much NH₄Cl will react to yield 500 g of NH₃?
18: 200 g of Na₂CrO₄ mixed with Pb(NO₃)₂ produce how many grams of PbCrO₄?
19: How much ZnSO₄ is formed by the reaction between H₂SO₄ and 500 g of Zn?
20: How much (NH₄)₂SO₄ is formed when ammonia reacts with 67.5 g of H₂SO₄?
21: What weight of water results from the reduction of 25 g of CuO?
22: How much water gas is made from 50 kg. of steam and carbon?
23: What weight of NH₄NO₃ will produce 200 g of N₂O?
WEIGHT - VOLUME PROBLEMS

1: 162 g of aluminum in dilute acid yields what volume of hydrogen?
2: 22.4 g of potassium reacts with water to produce how much H₂?
3: 18.37 g of KClO₃ are heated to yield what volume of O₂?
4: 250 g of CaCl₂ react with HNO₃ to give what volume of CO₂?
5: In the reaction between NH₄Cl and 230 g of NaNO₂, how much N₂ is formed?
6: 156 g of NaHSO₃ is treated with HCl to give what volume of SO₂?
7: How much H₂S can be obtained from 66 g of FeS?
8: If 132 g of NH₄Cl is heated with NaOH, how much NH₃ is produced?
9: How many liters of O₂ can be made from 530 g of KClO₃?
10: What weight of Zn will give 250 ml of H₂ when treated with acid?
11: How much SO₂ is formed when 100 g of Cu and 350 g conc. H₂SO₄ react?
12: How much O₂ is formed by the electrolysis of 5 kg. water?
13: What volume of CO₂ is prepared by burning 108 g of carbon?
14: What weight of Mg will release 124.4 liters of H₂ from an acid?
15: How much carbon is needed to yield 22.4 liters of CO₂?
16: What weight of H₂O decomposes to give 560 liters of O₂?
17: In a soda-acid fire extinguisher, how much soda is needed to produce 1,120 liters of CO₂?
18: How much iron is needed to give 92.4 liters of H₂ when reacted with steam?
19: H₂SO₄ solution weighs 143.2 g. How much does the solution weigh after 44.8 liters of NH₃ gas has been added to it?
20: How much ZnSO₄ is formed by the reaction between H₂SO₄ and 500 g Zn?
21: How much (NH₄)₂SO₄ is formed when ammonia reacts with 67.5 g H₂SO₄?
22: What weight of water results from the reduction of 25 g of CuO?
23: What weight of NH₄NO₃ will produce 200 g of N₂O?
24: How much CaCO₃ is needed to get 50 kg. CO₂?
25: How much SO₂ is formed by burning 70 g of sulfur?
NORMAL AND MOLAR SOLUTIONS

How many grams of solute are needed to prepare each of the following solutions?

1. 1 liter of 1.0 M NaCl
2. 5 liters 1.0 M Na₂CO₃
3. 6 liters 0.1 M H₂SO₄
4. 10 liters 0.2 M ZnCl₂
5. 500 ml 0.5 M KNO₃
6. 200 ml 0.03 M Ca(OH)₂
7. 250 ml 0.01 N H₂SO₄
8. 450 ml 5 M H₂SO₄
9. 200 ml 2M MgSO₄
10. 1 liter 0.1 N HCl
11. 2 liters 0.5 N H₂SO₄
12. 3 liters 0.3 H₃PO₄
13. 250 ml 0.01 N H₂SO₄
14. 3.5 liters 4.6N HNO₃
15. 42.5 ml 5N H₂SO₄
16. 34 ml 0.05 N Ba(OH)₂
17. 10 liters 6N KOH
18. 250 ml 0.05 N H₂SO₄
19. 250 ml 0.025 N Na₂CO₃
20. 750 ml 0.015 N CaCl₂
21. 1 liter 1.0 M Pb(NO₃)₂
22. What is the molarity of a solution containing 15.0 g K₂CrO₄ in 200 ml of solution?
23. What weight of (NH₄)₂SO₄ is required to prepare 400 ml of M/4 solution?
24. Determine the molarity and normality of a solution containing 4.055 g FeCl₃ in 50.0 ml of solution.
25. Calculate the normality of a solution containing 7.88 g HNO₃ per liter.
OXIDATION - REDUCTION REACTIONS

1: Cu + HNO₃ conc. → NO₂ + Cu(NO₃)₂ + H₂O
2: Cu + HNO₃ dil. → NO + Cu(NO₃)₂ + H₂O
3: KMnO₄ + H₂O + KI → MNO₂ + KOH + KIO₃
4: KClO₃ → KClO₄ + KCl
5: KMnO₄ + FeSO₄ + H₂SO₄ → K₂SO₄ + MnSO₄ + Fe₂(SO₄)₃ + H₂O
6: Cu + H₂SO₄ → CuSO₄ + SO₂ + H₂O
7: KMnO₄ + H₂SO₄ → K₂SO₄ + MnSO₄ + H₂SO₄ + H₂O
8: Ag + HNO₃ → AgNO₃ + H₂O + NO₂
9: CrI₃ + KOH + Cl₂ → K₂CrO₄ + KIO₄ + KCl + H₂O
10: H₂SO₄ + MnO₂ + KI → MnSO₄ + K₂SO₄ + I₂ + H₂O
11: MnO₂ + H₂SO₄ + NaBr → MnSO₄ + Na₂SO₄ + Br₂ + H₂O
12: KMnO₄ + H₂O + KBr → MnO₂ + KOH + KBrO₃
13: CuO + H₂ →
14: PbO₂ + H₂S → PbS + SO₂ + H₂O
15: K₂CrO₄ + HCl → KCl + CrCl₃ + H₂O + Cl₂
16: NaCrO₂ + NaClO + NaOH → NaCl + Na₂CrO₄ + H₂O
17: H₂SO₄ + O₂ →
18: H₂S + FeCl₂ →
19: PCl₅ + Cl₂ →
20: Cl₂ + NaI →
APPENDIX E
NEW ENGLAND ASSOCIATION OF CHEMISTRY TEACHERS

MINIMUM SYLLABUS FOR A COLLEGE PREPARATORY COURSE IN CHEMISTRY (1956)

The required section is divided into two parts, Descriptive Chemistry, Part I; and General Chemistry, Part II. The order of the topics as listed in the syllabus is not necessarily the order in which they should be taught. Every teacher must feel free to decide when and how to teach each topic.

The NEACT also stipulates that individual laboratory work, including preparation of gases, quantitative exercises, and ionic reactions should be an essential part of the course. At least one double period each week should be devoted to laboratory work and four single periods to classroom discussions and demonstrations.

Part III is a list of Supplementary Topics. It includes subjects that are important but not essential and is presented simply for the guidance of those teachers who have additional time or gifted students and who may select material from it adaptable to their time and course.

Part I. DESCRIPTIVE CHEMISTRY

A. Chemistry of some non-metals and their common compounds:
Oxygen, hydrogen, nitrogen (nitric acid and nitrates, ammonia, nitric oxide, nitrogen dioxide); sulfur (hydrogen sulfide, sulfur dioxide, sulfuric acid, and sulfates); carbon (carbon dioxide, carbon monoxide

\[ \text{Committee on the revision of the chemistry syllabus, NEACT, "A Minimum Syllabus for a College Preparatory Course in Chemistry", Journal of Chemical Education, (June, 1957) 34:6, pp. 307-308.} \]
carbonic acid and carbonates, methane, methyl- and ethyl alcohol; halogen family (halides and related acids)


5. Chemistry of three metals and their common compounds; sodium (sodium hydroxide, sodium chloride, sodium carbonate and sodium hydrogen carbonate); aluminum; iron.

   1. Definition and properties of metals in general
   2. Typical methods of extracting metals from their ores.
   3. Occurrence of three metals and their extraction from compounds and from ores.
   4. Simple reactions into which they enter.

6. Industrial Processes: Haber, Ostwald, Contact.

PART II. GENERAL CHEMISTRY

A. Physical Properties of solids, liquids and gases

   1. Kinetic-molecular theory; dependence of the kinetic energy notion of molecules upon temperature; differences between solids, liquids, and gases from kinetic-molecular viewpoint.

   2. Properties of solids and liquids.

   3. Properties of gases: effect of changes of temperature and pressure on the volume of gases; explanation in terms of the kinetic-molecular theory. Avogadro's Law; derivation of gram-molecular weight and volume. The
treatment of the gas laws is to be quantitative.

B. Chemical change.--

1. Elements, mixtures and compounds: nature of chemical changes; types of chemical reactions; electrochemical series.
2. Balancing of simple chemical reactions (by inspection only).
3. Quantitative relationships: problems based on chemical equations, by weight, by volume, and by weight and volume; derivation of formulas from percentage composition; mole and molar solutions.
4. Chemical combinations explained in terms of atomic structure: combination and replacement reactions in terms of electron transfer; ionic nature of substances entering double decomposition reactions; sharing of electron pairs in formation of covalent bonds.
5. Valence: ionic valence (electrovalence), equal to charge possessed by ion; emphasis on essential ionic nature of salts; electrolysis of fused salts; covalence in simple molecules, equal to number of pairs of electrons shared with other atoms; emphasis on essential non-ionic structure of non-electrolytes.

C. Solutions.--

1. Definition of concentrated and dilute solutions.
2. Solutions of electrolytes; complete dissociation
of salts, strong acids and hydroxides of alkali metals; degree of ionization of weak acids and weak bases (\(\text{H}_2\text{O} \cdot \text{OH}\));
definition of acids and bases (Arrhenius concept); hydrolysis of salts – reactions of ions with water; reason why some ionic reactions go to completion (formation of weakly ionized product, and formation of an insoluble gas or solid); electrolysis of aqueous solutions.

D. Structure of matter.—

1. Definition of atom and molecule; Dalton's Atomic Theory.
2. Nuclear charge and the arrangement of electrons (limited to the first twenty elements)
3. Periodic Law and its relationship to atomic structure (brief treatment only); elements should be studied with reference to their position on table)
4. Explanation of periodicity in terms of properties of elements.

E. Nucleonics.— Radioactivity, isotopes, agents of transmutation (alpha particles and neutrons); fission; fusion.

PART III. SUPPLEMENTARY TOPICS

A. Organic chemistry.— simple hydrocarbons, alcohols, acids, esters, aldehydes, ketones, isomerism.
C. Molecular weights.— Determination by depression of freezing point and elevation of boiling point.
D. Equivalent weights.— Normal solutions.
E. Balancing equations by electron transfer.— Oxidation-reduction.
F. Metals.— copper, zinc, lead, titanium.
G. Non-metals.— phosphorus, silicon.
H. Applied chemistry.— plastics, rubber, glass, cement, textiles, and foods.

The Committee on the Revision of the Chemistry Syllabus which reviewed the minimum syllabus was comprised of the following:

Maurice M. Whitten, Chairman, Gorham State Teachers College, Gorham, Maine.

Sister Ernestine Marie, Magr. Ryan Memorial High School, Dorchester, Massachusetts.

Anna Jane Harrison, Mount Holyoke College, South Hadley, Massachusetts.


Dorothy N. Gifford, The Lincoln School, Providence, Rhode Island.
APPENDIX F

PACKET SENT TO SCHOOLS PARTICIPATING IN PHASE II
Teacher's Check List

You will find the package we have sent you includes both inductive and deductive class notes and laboratory directions. If you have two classes in chemistry in your school, will you use the two methods, one on each? If you do not have two classes, could you find a teacher of your acquaintance to follow the deductive notes in his or her class this fall? The deductive section is, for the most part, very much like what most teachers teach anyway. We would like to have pairs of classes for our records, one inductive and one deductive, both of which have taken the same pre- and post-test.

1. Administrator the pre-test and send the papers to us for scoring. You may hold the entire set until the post-tests are also ready. If you wish, be sure all papers are clearly marked with the first name and last initial of the pupil.

2. Check with the office to get a list of those families not here so that you can send a report card in your own hand. They may be very interested to hear how well the tests turned out.
There are spaces provided for you to make comments throughout the papers. You may fill in any of these that you wish.

After the unit is taught, administer the post-test and send the papers to me for scoring. Again, make sure that each paper is clearly marked with the first name and last initial of the pupil.

In the final package you are sending back to me there should be:

1. Pre-tests (for inductive and deductive classes)
2. Intelligence test scores (for inductive and deductive classes)
3. School data sheet (for each class used)
4. Inductive unit sheets with comments; deductive evaluation sheets
5. Inductive laboratory teacher's sheets with comments
6. Post-tests (for inductive and deductive classes)

You may retain any other material, including the lecture notes, audio notes, or deductive teaching and laboratory notes.

Please refrain from using any Audiovisual aids until the post-tests are back from me.
The teaching of the required course of chemistry in Grade XII or XI in the college preparatory course in high school must, because of the impending CEEB examinations and the need of fulfilling college entrance requirements, be formal. Such formality can have a definitely stultifying and restricting effect on vital science and in many instances causes the subject-matter mastery to deteriorate into mere parrot-like memorization.

It is difficult to capture a feeling of scientific zest in such a topic as equation-balancing, which is probably the most prosaic and mechanical of an alarmingly dull list of requirements. However, this unit can come alive if the teacher will but try to give it meaning. It need not be a feat of sheer memorization; it can and should be a study in understanding and problem-solving.

**PUPIL-TEACHER GOALS**

1. The student should learn CONCEPTS through organized thinking and the concept that is being applied here is the one of equation-balancing.

2. The student should develop some PROBLEM-SOLVING ABILITY as in the preparation of salts by various methods.

3. The student should learn those particular SKILLS which are specific to chemical laboratory work in testing the reactivity of substances.
The unit should make laboratory work particularly interesting and should increase observational powers and call attention to hidden meanings to be discovered in laboratory experimentation.

Some students may see the practicality, versatility, and great future of chemistry, and be interested enough to make chemistry a vocational choice.

A greater appreciation of industrial chemistry and its impact on modern living should result.

A better understanding of the economy of the nation and its relation to chemistry will result from:

a: readings in current periodicals.
b: study of industrial processes, particularly of gaseous reactions.
c: understanding of the Haber process for manufacturing ammonia.
d: study of hydrocarbon combustions.

Core Activities

1. Read the assigned chapters on equation-balancing and reaction classification, solubility, and equilibrium, in your textbook.

2. Listen to lectures given by the instructor on:
   a: laws of conservation of matter and of energy
   b: meaning of equations
   c: equilibrium: law of mass action
   d: law of le chatelier

3. Do laboratory work without detailed and specific instructions. Ideas and suggestions may be gleaned from outside reading.
   a: EMF Series
   b: synthesis and analysis
   c: single and double replacement
   d: reactions that go to completion
   e: four ways to prepare a salt

**RELATED ACTIVITIES**

Read a description of the Haber Process and apply the Law of Mass Action and Le Chatelier's Principle to it.

<table>
<thead>
<tr>
<th>Those who would like to do this</th>
<th>SIGN HERE!</th>
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<tbody>
<tr>
<td>Name</td>
<td>Date Finished</td>
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<tr>
<td>Sally J.</td>
<td>3/2/52</td>
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</table>

Study the common ion effect in a college chemistry text such as *Qualitative Analysis and Chemical Equilibrium* by Hognes and Johnson.

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<td>Jamie P.</td>
<td>10/4/52</td>
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</table>

Visit a chemical laboratory and report on some reaction they were using or investigating.

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<tr>
<th>Name</th>
<th>Place</th>
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<tbody>
<tr>
<td>Charles S.</td>
<td>Lab. Spec. Work</td>
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**Field trips?**

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<td></td>
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</table>
Read the life of some great chemist and give an oral report on his contribution to chemistry.

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<tbody>
<tr>
<td>John H. Priestly</td>
<td>4/15/57</td>
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Perform a quantitative analysis or synthesis reaction and write up your work.

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Read the chapters on equation-balancing, equilibrium, reaction classification and solubility in two other text books.

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Enjoy thorough study?

SIGN UP!
Here is a way in which any good teachers can see that their students are acquiring some of the skills and behaviors which are considered desirable. Will you make records of pupils responses to the following headings?

**V2** The skill and knowledge shown in laboratory work, aside from any mark on the accuracy of testing and identification, can evaluate general ability and intelligence.

*Sample:*

Dollie B. keeps her work areaspitless; George B. has broken 7 test-tubes and 3 flasks so far!

Please fill in your comments on the page facing this one.

**V3** The interest shown in laboratory work, the attentiveness, the observations, questions, inventiveness, all will indicate the students' approach to a scientific attitude.

*Sample:*

Opposition from the left to try out experiment with Andy. Almost invited to help next.

**V3** The problem-solving part of a unit in which students must devise, carry out and explain laboratory procedures will tell much about their ability to reason.

*Sample:*

John plans to have detailed notebook with detailed observations, conclusions & references.

**V4** The chemicals and equipment the student requests are a good indication of his growth in and assimilation of chemical knowledge.

*Sample:*

Roger F. requested additional metal to further test the O.F.M.F. series.
Use of reference in correlation with observations indicates an ability to synthesize the material gathered in reference to the point under investigation.

Example:

George P turned to the "Handbook of Chemistry and Physics" to determine the relative solubilities of two cadmium salts.

The type of reference book used and the student's interpretation of it is a clear indication of his level of learning.

Example:

Hayden's used Robert's "Handbook" as a related activity and textbook. John's used a high school lab manual for his.

The final report sheets will show the degree to which the material has been assimilated.

Example:

James's prepared in unity a unorganized report sheet.

The procedures used in the laboratory will indicate and differentiate between the leaders and the followers, the thinkers and the technicians.

Example:

Betty was to weigh 5 g on an electronic. Then add and indicate it.

The approach to the laboratory problem of preparing a salt in four ways will indicate the amount of understanding the student has of what has gone before.

Example:


Please report these on the opposite page, and insert more pages if you wish.
Equation-balancing through laboratory work

The unit on chemical equation balancing is designed to elicit student understanding and to enlarge student concepts of scientific principles and processes. Equations can be balanced, all too frequently, in a purely mechanical way, without knowledge of the meaning. What are coefficients? What do they signify? Of what use are they? What connection have the Law of Conservation of Matter and the Law of Conservation of Energy with equation balancing? In replacement reactions, how do you know what replaces what? How do you know what reactions occur and what the products are? These questions are often asked by students, and more often arise but are not voiced. The unit here presented is designed to answer these and similar questions by introducing students to the operations of equation balancing by way of the laboratory. The students will see the reality before they learn the symbolic representation of it. They will work from the concrete to the abstract and in this way should gain a firmer and more certain grasp of the whole problem of equation balancing.

Keep a diary

We encourage, and in fact, urge you to make any remarks you may care to write. You will find space provided in several places for this writing; you are always at perfect liberty to add more sheets of comments anywhere you choose. We would like to have this unit returned to us with a running commentary — a diary of your experience in teaching it. On activity pages, on class and laboratory notes, you will find columns, blocks, and boxes — all designed for your remarks, your date and time schedule for that particular section of the teaching. We are also asking for intelligence test results on the pupils being used in this study, as they are very important to us. If, for any reason, your school authorities do not wish to reveal this information, send us your data anyway; they will still be of great value!

The sheets on "related activities" have boxes left for student signatures. We suggest that you pass these sheets around for students to read and to sign for the activities they intend to perform. The date the assignment is completed you may record.

You will notice that many sheets are stamped "return this". These papers we want back in the final package with the answer sheets for us to score.

We hope that you will enjoy working this unit with us, and we wish you the best of luck with your classes this year!
The fundamental laws of conservation easily occupy a class period. First, a discussion of burning wood -- would the products weigh the same as, more than, or less than the wood? Then all the reactants? Is matter lost? Is energy lost or created? A round-the-room discussion of this nature will open the field. Then the concept seems quite clear, a final statement of the Law may be given: Neither matter nor energy can be created or destroyed in a chemical reaction.

A slight digression into atomic reaction may be made with mention of the equivalence of matter and energy from the Einsteinian viewpoint. The equivalence formula \( E = mc^2 \) should be used.

Recast the work in the form of some word equations:

- Wood + oxygen produces ashes + gases + heat.
- 50 grams of wood and oxygen produce 50 grams of ash and gas.
- Pure carbon + oxygen → carbon dioxide; and again, 10 grams of carbon and oxygen will yield 10 grams of carbon dioxide.

-2-

SYMBOLIC EQUATIONS

All students have presumably made the acquaintance of the Periodic Table and are familiar with the periodic classification of the elements and with the common valence states. They can write the formulas for compounds correctly and know which gases are diatomic.

Suggest that the equation on carbon mentioned before be written symbolically: \( C + O_2 \rightarrow CO_2 \)

Check by counting the number of atoms on the left and right of the arrow. \( 1 C \rightarrow 1 C \)

\( 2 O \rightarrow 2 O \), therefore, all is in accord with the Law of Conservation. How about the decomposition of water? Set up a demonstration and let it run until the tubes contain enough gas for the differences in volume to be clearly discernable. Any type of electrolysis apparatus you have or can make will do the work satisfactorily, I'm sure. \( H_2O \rightarrow H_2 + O_2 \), but upon counting the atoms we find \( 2 H \rightarrow 2 H \) and \( 1 O \rightarrow 2 O \), hence there is a discrepancy and the Law of Conservation is broken. We have gained matter in our reaction. If we take twice as much of everything, will it help?
Evidently we have merely doubled everything and are no better off. Meanwhile the demonstration has been running and can now be viewed. (Check your equipment and see the time needed. It may have to be set up an hour before class be-ins.) What does our demonstration suggest to us? Re: the cathode and anode volumes, identify the gases at each pole, and write the obvious reaction they dictate:

\[ 2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2 \]

4H \rightarrow 4H

20 \rightarrow 2 0, and here we finally have satisfaction of the law. Point out then that the volumes of gaseous substances that react are represented by the coefficients of the molecules in the balanced reaction.

Just what do the coefficients signify? They tell us the number of molecules, but is that all? In the case of gases we have just seen they represent the volume of the gases reacting.

2 vol. of steam, containing the same number of molecules and the same total number of atoms as in products

1 volume of oxygen containing the same number of molecules

2 volumes of hydrogen each containing the same number of hydrogen molecules

The coefficients represent the relative amounts of the reactants and products; for gases they signify volumes; for solids they indicate combining weights.

We have seen that \( 2 \text{H}_2 + \text{O}_2 \) is formed from \( \text{H}_2\text{O} \). In other words, hydrogen combines with oxygen in a fixed ratio and that ratio is the ratio of the coefficients in the balanced equation, viz. 2:1. If we have equal amounts of hydrogen and oxygen given us, what will happen? (Elicit student response)

In the case of solids the same things happens:

\[ \text{Fe} + \text{S} \rightarrow \text{FeS} \]

This is a 1:1 ratio which means 1 molecule of iron will combine with one molecule of sulfur, or one gram molecular weight of iron will combine with one gram molecular weight of sulfur, or 56 grams of iron will combine with 32 grams of sulfur and will produce 88 grams of FeS.
Suppose I have 50 grams of sulfur given me? How much will remain in excess? (18 grams according to the Law of Conservation)

Our coefficients tell us what relative weights or volumes can be made to combine and the relative amount of product we will get by the combination. No equation is balanced until the atoms of products and reactants are equal in number.

Remember -- never change the subscripts of a compound to balance the equation. Use only variations of coefficients, of amounts of substances.

In the laboratory you will do combinations, analyses, and single replacement and double replacement reactions -- all are balanced in accord with the same basic Law of Conservation.

-3-

(EQUILIBRIUM)

(120 minutes)

In the laboratory you have found that solutions of soluble compounds that produce soluble substances do not apparently react.

\[
\begin{align*}
\text{NaCl} + \text{KNO}_3 & \rightarrow \ ???? \\
\text{KCl} + \text{NaNO}_3 & \rightarrow \ ????
\end{align*}
\]

Nothing visible happens in the cool test tube in either case. We could write the ionic equation:

\[
\text{Na}^+ + \text{Cl}^- + \text{K}^+ + \text{NO}_3^- \rightleftharpoons \text{K}^+ + \text{Cl}^- + \text{Na}^+ + \text{NO}_3^-
\]

We find the very same ions on both sides of the equation and since all four ions are free in the solution, we should expect little in the way of a complete reaction. We say we have a reversible reaction. It is not a state of stagnation but of dynamic equilibrium as these ions are momentarily linked up with partners of the opposite charge, and then break away again and link with another partner, and so on.

Let us consider the gaseous equilibrium which is involved in an important industrial process:

\[
\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2 \text{NH}_3
\]

This is the Haber process for the fixation of nitrogen by the synthesis of ammonia. All products are gaseous. Looking at the coefficients we have: 1 volume of \( \text{N}_2 \) will combine with 3 vol. of \( \text{H}_2 \) and will produce 2 vol. of ammonia gas. But it is relatively easy to decompose ammonia into hydrogen and nitrogen:

\[
2 \text{NH}_3 \rightleftharpoons \text{N}_2 + 3 \text{H}_2
\]

Both forward and reverse reactions can occur. Which will occur if we carry on the reaction in an enclosed system? Both will. We will have a dynamic equilibrium with nitro en and hydro en combining to make u-monia and ammonia decomposing to form nitrogen and hydrogen.
Which will go faster? That all depends on the external circumstances we impose. What variables can we have?

1: concentration
2: pressure
3: temperature

Let's consider concentration first. Suppose we separate the NH₃ as soon as it is made and remove it from the field of action. We no longer can have ammonia decomposing to give back the original atoms of nitrogen and hydrogen from which it was made. It is safely cooled and stored in liquid form in pressurized containers. We have broken the equilibrium and have permitted the forward reaction to go to completion. The forward reaction has won.

Let's upset the equilibrium again, this time by adding extra hydrogen to the reaction above in the closed reaction chamber. Momentarily things are awry as we have too much hydrogen for the equilibrium condition and we must get it used up. How? By combining it with N₂, so the reaction speeds up in the forward direction to use up this excess hydrogen. Meantime, extra NH₃ is created, which is a further upset to the original equilibrium condition. It will have to decompose to give more N₂ and H₂; hence, that reverse reaction is speeded up also for a while, and finally all events out and we have once more reactions proceeding in different directions at the same rate.

This whole concept is embodied in the Law of Mass Action: The speed of a chemical reaction is proportional to the products of the concentrations of the reacting substances. (This material will have to be reviewed with student participation and illustration with different equilibria such as: \(2 \text{CO}_2 \rightleftharpoons 2 \text{CO} + \text{O}_2\) and \(\text{CH}_4 + \text{O}_2 \rightleftharpoons \text{CO}_2 + 2 \text{H}_2\text{O}\).

\[= 4 =\]

**LE CHATELIER'S PRINCIPLE**

Last time we considered the effects of volume changes on a system in equilibrium. Now let us take cognizance of temperature and pressure changes. Gases are very susceptible to pressure and temperature variations so let us turn back to the study of the Haber reaction:

\[\text{N}_2 + 3 \text{H}_2 \rightleftharpoons 2 \text{NH}_3\]

High temperature favors the dissociation of molecules into simplest forms as internal agitation increases beyond the bond strength of the molecule and it can no longer hold together. Hence, the increased temperature will cause the backward reaction to proceed with greater speed. Also, high temperatures favor large volumes as heat causes expansion, so too, the backward reaction toward the larger volume of
reactants will be favored.

What about pressure? High pressures tend to create small volumes as pressure compresses. In our reaction we have 4 volumes of gas on the left side and 2 on the right. Hence, pressure will hasten the forward reaction producing the smaller volume. This can be demonstrated by an irregular balloon or beach toy of the rubber whale variety, plus a board. Pressure on the high side will cause the molecules of gas in the toy to move to the smaller end.

Discussion about other gaseous reactions will be needed. First, call for student balancing of the reaction, and then for pressure and temperature explanations. Volume and concentration changes should also be reviewed.

Some students may introduce the question of catalysts and their effect on equilibrium shift. The underlying discussion is a bit too far afield for the regular high school student, but it can be pointed out and demonstrated that doubling the amount of catalyst does not double product volume. An extra assignment may be done by those interested in perusing the literature on this topic.
<table>
<thead>
<tr>
<th>Equation</th>
<th>Reaction</th>
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<tbody>
<tr>
<td>1: C + O₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>2: Fe + O₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>3: P + O₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>4: Na₂ + H₂</td>
<td>( \rightarrow )</td>
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<tr>
<td>5: Al + NaOH</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>6: Fe + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>7: CH₄ + O₂</td>
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<tr>
<td>8: KClO₃</td>
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<tr>
<td>9: Cu + H₂SO₄ conc.</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>10: Na₂SO₃ + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>11: CaCO₃ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>12: AgNO₃ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>13: BaCl₂ + Na₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>14: K₂CrO₇ + PbCl₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>15: As(NO₃)₂ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>16: ZnCO₃ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>17: PbSO₄ + (NH₄)₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>18: CaO + C</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>19: H₂O electrolysis</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>20: Mg + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>21: Fe₂O₃ + Al</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>22: AgNO₃ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>23: CuO + H₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>24: FeS₀₃ + C</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>25: K₂Fe(CN)₆ + FeCl₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>26: K₂CO₃ + PbCl₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>27: K + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>28: Al₃(SO₄)₃ + NH₄OH</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>29: (NH₄)₂SO₄ + KOH</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>30: NaNO₃ + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>31: MnO₂ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>32: Cu + Cl₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>33: SbCl₅ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>34: Pb(NO₃)₂ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>35: CaO₂ + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>36: Ca₂O₃ + Al</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>37: Cu(OH)₂ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>38: NaI + Br₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>39: Na₂CO₃ + AgNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>40: NaCl + AgNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>41: Sb + Cl₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>42: H₂SO₄ dil. + Fe</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>43: Al + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>44: CaCO₃ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>45: H₂SO₄ + Ca(OH)₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>46: H₂SO₄ + NaOH</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>47: Na₂CO₃ + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>48: CaO + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>49: NaCl + HNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>50: Fe₃O₄ + CO</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>51: K + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>52: Ca + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>53: CaSO₄ + Na₂CO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>54: NaNO₃ + KCl (heated)</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>55: MgBr₂ + Cl₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>56: NaHCO₃ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>57: KHSO₃ + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>58: Na₂H₃O₂ + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>59: KCl + H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>60: Cu + conc. HNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>61: Cu + conc. H₂SO₄</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>62: Fe₂O₃ + Al</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>63: Cu + dil. HNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>64: Ca(OH)₂ + HNO₃</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>65: Al + O₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>66: CuSO₄ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>67: Zn(NO₃)₂ + H₂S</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>68: H₂O heated</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>69: CaCO₃ + H₂O₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>70: Ca(HCO₃)₂</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>71: Na₃SO₄ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>72: NH₄HCO₃ + NaCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>73: NH₃ + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>74: NH₃ + HCl</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>75: SO₂ + H₂O</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>TABLE OF COMMON VALENCES</td>
<td></td>
</tr>
<tr>
<td>---------------------------</td>
<td>--</td>
</tr>
<tr>
<td>ammonium NH₄⁺</td>
<td>barium Ba⁺⁺</td>
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<tr>
<td>hydrogen H⁺</td>
<td>calcium Ca⁺⁺</td>
</tr>
<tr>
<td>potassium K⁺</td>
<td>cupric Cu⁺⁺</td>
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<tr>
<td>silver Ag⁺</td>
<td>ferrous Fe⁺⁺</td>
</tr>
<tr>
<td>sodium Na⁺</td>
<td>lead Pb⁺⁺</td>
</tr>
<tr>
<td>cuprous Cu⁺</td>
<td>magnesium Mg⁺⁺</td>
</tr>
<tr>
<td>mercurosus Hg⁺</td>
<td>mercuric Hg⁺⁺</td>
</tr>
<tr>
<td>zinc Zn⁺⁺</td>
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</tbody>
</table>

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<thead>
<tr>
<th>TABL</th>
<th></th>
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</thead>
<tbody>
<tr>
<td>acetate C₆H₅O₂⁻</td>
<td>carbonate CO₃²⁻</td>
<td>phosphate PO₄³⁻</td>
<td></td>
</tr>
<tr>
<td>bicarbonate HCO₃⁻</td>
<td>oxide O⁻</td>
<td>phosphite PO₃⁻</td>
<td></td>
</tr>
<tr>
<td>bisulfate HSO₄⁻</td>
<td>sulfate SO₄²⁻</td>
<td></td>
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<tr>
<td>chlorate ClO₃⁻</td>
<td>sulfite SO₃⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>chloride Cl⁻</td>
<td>sulfide S⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydroxide OH⁻</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>iodide I⁻</td>
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<tr>
<td>nitrate NO₃⁻</td>
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<td></td>
<td></td>
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<tr>
<td>nitrite NO₂⁻</td>
<td></td>
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</table>

<table>
<thead>
<tr>
<th>ELECTROMOTIVE SERIES</th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>cesium</td>
<td>potassium</td>
<td>aluminum</td>
<td></td>
</tr>
<tr>
<td>sodium</td>
<td>calcium</td>
<td>zinc</td>
<td></td>
</tr>
<tr>
<td>magnesium</td>
<td>chloride</td>
<td>iron</td>
<td></td>
</tr>
<tr>
<td>iron</td>
<td>lead</td>
<td>tin</td>
<td></td>
</tr>
<tr>
<td>lead</td>
<td>hydrogen</td>
<td>mercury</td>
<td></td>
</tr>
<tr>
<td>hydrogen</td>
<td>copper</td>
<td>silver</td>
<td></td>
</tr>
<tr>
<td>copper</td>
<td>gold</td>
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</tbody>
</table>
This section of the laboratory work should precede any class discussion concerning the balancing of equations because it is the foundation for the understanding of the meaning of reactions.

A student is supplied with magnesium ribbon, red phosphorus, powdered and lump charcoal, iron filings, powdered and roll sulfur.

By burning a piece of magnesium ribbon the student sees immediately that products and reactants are not the same things. It may take much teacher questioning to bring out the nature of this change, and to convince the student that oxygen is the only other substance involved. However, finally the student can put on paper:

\[ \text{Mg} + \text{O}_2 \rightarrow \text{MgO} \]

It may be well to leave this equation unbalanced at first unless the student himself brings up the question. If he has studied the law of Conservation of Matter in class he should be puzzled by such a solution, but if he does not seem to notice the lack of balance, do not force it upon his attention now.

Let him take roll sulfur and burn it. Again his senses tell him the product differs from the reactants, and more quickly he writes:

\[ \text{S} + \text{O}_2 \rightarrow \text{SO}_2 \]

Carbon may be used in like manner, and a drop of limewater suspended from the tip of a stirring rod held in the mouth of the tube in which the charcoal powder is burning, or above the spatula-tip on which it is ignited, will give evidence for the presence of \( \text{CO}_2 \). This latter test the student should find for himself in his text or some other reference book in the laboratory.

Phosphorus burned on a spatula-tip will give another product noticeably different from the reactants. The student may write either:

\[ \text{P} + \text{O}_2 \rightarrow \text{PO}_3 \]

or

\[ \text{P} + \text{O}_2 \rightarrow \text{PO}_3 \]

The order in which the above reactions are done is entirely a free choice. The student may select the chemicals he wishes in any order, but at least this many should be handled to give him a fair comprehension of the distinctive powers of atmospheric oxygen and the cooperative non-reactivity of other atmospheric components.
When these many reactions above have been completed the teacher should suggest that all reactions written in the symbolic form should be checked to see that the atoms on each side of the reaction arrow are alike in number.

75 minutes

ANALYSIS

Date and time:
Date and time:

The idea of performing the reverse of synthesis may now be introduced by placing bottles of mercuric oxide and potassium chlorate on the table for student use. Most of the students will heat these compounds in a tube as they did in the synthesis experiments, or they may try heating them on the tip of a spatula. If so, encourage the use of the tube. If mercuric oxide is thus heated, the student will find globules of mercury on the walls of the tube and a gas escaping. The teacher may insert a glowing splint into this gas, or, if the student has done the experiment on the preparation of oxygen he should be led to make this test himself. With the evidence thus accumulated the student can then write:

\[ \text{HgO} \rightarrow \text{Hg} + \text{O}_2 \]

He may then be left to try other oxides (after he has balanced the HgO equation) and may attempt water. If he does, his tests for O\(_2\) and H\(_2\) proving negative, he will see for himself that HgO is not an easily decomposed oxide.

KClO\(_3\) heated in small quantities is relatively safe and will give an O\(_2\) test. If the oxygen experiment has been done and he has understood it, he will request MnO\(_2\) as a catalyst. So much the better as he has certainly grasped the idea of catalysts and catalysis. Even if the oxygen experiment has not been done, the teacher may feel safer with the students using the MnO\(_2\) with the KClO\(_3\) and may suggest its use. KClO\(_3\) being a ternary compound introduces a complication to the student: does he have 2 or 3 products? He may want to write:

\[ \text{KClO}_3 \rightarrow \text{K} + \text{Cl} + \text{O}_2 \] (or even O\(_3\))

Attention should then be called to the difference in structure between atoms and molecules and ions, and the student should consult a text for tests for the identification of Cl\(^{-}\), Cl\(_2\), K\(^{+}\) and K. He will find the ionic tests easy to apply and will doubtless perform these first, thus finding out that
he does have K⁺ and Cl⁻, hence he can then write:

\[ \text{KCl} \rightarrow \text{K}^+ + \text{Cl}^- \]

The mechanics of balancing the number of atoms in the left and right members of this equation can be done in the laboratory or in a class session in connection with the Law of Conservation of Matter.

It is suggested that the student now turn to reactions between other elements such as powdered sulfur and iron filings. This reaction is best carried out in a crucible as the FeS formed can be easily removed from this vessel whereas it destroys a test tube. If the student has grasped the basis of the synthesis type of reaction he will immediately write:

\[ \text{Fe} + \text{S} \rightarrow \text{FeS} \]

And now he has a broader concept of synthesis—simple substances combining, usually with the aid of heat, to form more complex substances. Here other possible reactions may be suggested by the teacher, or devised by the students, the latter being the preferable course. Cu with O₂ or S, Ag with O₂ or S, and such combinations may be carried out in simple fashion.

---

**75 minutes**

**ELECTROMOTIVE SERIES**

Date and time:

Date and time:

The directions given the students are minimal, hence, each student must work out the methods of testing for himself. In order to prevent completely unrelated and unnecessary experimentation the materials to be used are listed. Provided a student has made use of these and then still wants other chemicals, it is generally a sign he has grasped the significance of the experiment and would like to investigate some point even further. If, however, he has not used the prescribed material and wants other substances, his request is often a bid for mere time-wasting work.

Students are warned to be careful with calcium and sodium on to use only minute quantities of these metals with water. Advise them to wear goggles when working with alkali metals. Their knowledge of the Periodic Table should guide them in using other dangerously reactive metals. "Test the products of the reactions," "Test all products" appears twice in the short guide sheets to impress this point—the student knows the reactants and he must proceed to discover positively the products he has so that he can write the proper symbols on both sides of the chemical equation. He
puts sodium with water and identifies hydrogen gas by means of a burning splint, OH⁻ by litmus or other indicator, and Na⁺ by a flame test. He can then write with conviction and understanding:

\[
Na + H_2O \rightarrow H_2 + NaOH
\]

If he has an understanding of the basic concepts of conservation of matter which he should have grasped from lectures, along with a knowledge of the diatomic nature of hydrogen gas, he will see immediately that he has more hydrogen as a product than went into his reaction as reagent. Here he may need help, or he may be able to work out the problem for himself. If this experiment is performed after the basic concepts of equation-balancing have been taken up in lecture, then he should work it out by himself.

In other laboratory testing he will add zinc to dilute sulfuric acid and identify hydrogen gas by a burning splint, zinc ions by hydrogen sulfide precipitation and sulfate ions by barium sulfate precipitation. He can then write:

\[
Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2
\]

But in addition he will also know that

\[
BaCl_2 + ZnSO_4 \rightarrow BaSO_4 + ZnCl_2
\]

Moreover, he sees that ZnSO₄ is soluble whereas BaSO₄ is not, so that in this way solubilities of common chemicals will become part of his fund of ready knowledge.

He has also, probably, discovered that small quantities of reagents can be handled easily and give him neat, clean, quick reactions. By trial and error he knows the value of clean glassware which will not contaminate tests. He is gaining facility in handling stock and reagent bottles and in manipulating his own glassware and burner.

He works with concentrated and dilute acids and finds that the concentrated ones are not quicker reactors, as he might have thought they were. This puzzle he will probably be unable to solve yet, but he has developed a mental set, ready and receptive to the coming section on ionization.

He finds that HNO₃ and concentrated H₂SO₄ do not behave like other acids with zinc and lead and other common metals. His burning splint test fails to identify the gaseous product. Perhaps he cannot write the equation involving these acids, but will have to refer to a text to find out the products that might be expected, and the tests for them. Very well, he has learned an important point that hours of classroom drill generally fail to impart. And if he consults a text to solve his problem, he has developed a sense of mastery — and a respect for and familiarity with his textbook.
In using various metals with HCl to determine the order of reactivity, he may get strange results and discrepancies unless the instructor has comparable forms of all the metals. If all are in strip, or all in pellets, or all in powder of the same mesh, this won't happen; but if the zinc is mussy, Cu in strip, aluminum is shot, and Zn in powder, the student may have a few too many variables for one experiment.

75 minutes

**SINGLE REPLACEMENT REACTIONS**

The EMF series of experiments has familiarized the student with single replacement actions so that there should be no need of repetition of that work.

Attention might well be called, however, to the gas formation common to so many of the single replacement reactions, and the precipitate formation common to so many of the double replacement reactions. This will pave the way for the work on reactions that go to completion.

**REACTIONS THAT GO TO COMPLETION**

Recall evidence which has been accumulated concerning reactions products and note that precipitate formation or gas evolution is apparently indicative of reaction. A different product is formed, but in order to be sure we have a product, some substance must be removed from the field of action. A gas removes ions that have combined into atoms or molecules; a precipitate removes ions that have combined into insoluble salts. Any un-ionized or very slightly ionized substance will perform the same role and remove ions from the field of activity, thus forcing reactions to completion and preventing reversibility.

An NaOH-HCl neutralization might be suggested as the products he e should be identifiable, after evaporation of the H₂O₂, both by taste and by crystal structure. Much chalkboard illustration of reversible and non-reversible reactions may be needed with some classes at this point. They should be encouraged to set up various solutions and see what evidence they obtain for reaction and then report this reaction as complete or incomplete, on the basis of their positive evidence.

Date and time:

Date and time:
Solutions of common salts should be available and the students should be encouraged to pair them up, starting with the ones they have used before.

In the previous experiment they discovered a test for a chloride, viz. the addition of AgNO$_3$ to a soluble chloride, hence, they can take all available chlorides and add to them AgNO$_3$, thus obtaining a precipitate of AgCl.

\[
\begin{align*}
\text{NaCl} + \text{AgNO}_3 & \rightarrow \text{AgCl} + \text{NaNO}_3 \\
\text{KCl} + \text{AgNO}_3 & \rightarrow \text{AgCl} + \text{KNO}_3 \\
\ast \text{CaCl}_2 + \text{AgNO}_3 & \rightarrow \text{AgCl} + \text{Ca(NO}_3)_2 \\
\ast \text{BaCl}_2 + \text{AgNO}_3 & \rightarrow \text{AgCl} + \text{Ba(NO}_3)_2 \text{ etc.}
\end{align*}
\]

Attention must be called to all unbalanced equations which violate the Law of Conservation of Matter and the student must go back over these and finish them if he has written them in an incomplete form.

By using solutions of the chlorides he should realize that most chlorides are soluble; similarly, he obtains nitrates and finds that they are likewise soluble. AgCl is the only insoluble salt he has probably found so far.

He may now work with sulfates and add to them BaCl$_2$ and derive a series of insoluble sulfates of barium, again learning the solubility of sulfates except that of barium in particular.

He may now work with miscellaneous pairs of solutions and if he takes such combinations as NaCl + KNO$_3$ he will find no evidence of action and hence he must write either

\[
\begin{align*}
\text{NaCl} + \text{KNO}_3 & \rightarrow \text{N.R.} \\
\text{or} \\
\text{NaCl} + \text{KNO}_3 & \rightarrow \text{KCl} + \text{NaNO}_3
\end{align*}
\]

In the case of Na$_2$SO$_4$ and BaCl$_2$, does he know which product is the precipitate? From his work on chlorides and from he should realize that NaCl is soluble and from his testing of SO$_4$ he should realize that BaSO$_4$ is not; does he?

(Remarks thus far.)
In the course of the work he will obtain several precipitates, probably PbCrO₄, PbCl₂, AgBr, BaCO₃ among others. At this time he may be introduced to a standard chart of solubilities after he has made a table of his own in regard to the relative solubilities of the compounds he has met. Na⁺, K⁺, and NH₄⁺ should at least have impressed him with their solubility.

There is a chart of solubilities included in this unit for your convenience.

At the end of this laboratory session the student should have a grasp of double decomposition reactions, and of the idea of differential solubility.

Remarks:

75 minutes

FOUR WAYS OF PREPARING A SALT

Date and time:

Now, can the student use the information he has acquired thus far? Can he make the practical application in preparing a pure salt by each of the four different methods: direct combination (synthesis), neutralization (double replacement), formation of a gaseous product (single or double replacement) and formation of a precipitate (double replacement)?

If he comes in with a method similar to the following, he has certainly mastered the topic:

\[
\begin{align*}
2 \text{Na} + \text{Cl}_2 & \rightarrow 2 \text{NaCl} \\
\text{NaOH} + \text{HCl} & \rightarrow \text{NaCl} + \text{H}_2\text{O} \\
\text{Na}_2\text{CO}_3 + 2 \text{HCl} & \rightarrow 2 \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2 \\
\text{Na}_2\text{SO}_4 + \text{BaCl}_2 & \rightarrow 2 \text{NaCl} + \text{BaSO}_4
\end{align*}
\]

Teachers may prefer to omit the direct combination reaction in actual experimentation and should feel free to do so. Those laboratories using semimicro procedures would be perfectly safe in performing it.

Some students may pick 4 different salts, and prepare one by each method, and while not so thought provocative, it will do, provided it is not a mere repetition of the experiments already done such as NaCl + AgNO₃; NaOH + HCl; CaSO₄ + BaCl₂ etc. This is hardly anything but parrot work and does not signify mastery of principle.

Insist on seeing student plans for this experiment before letting them proceed with the laboratory work.
1: \( \text{C} + \text{O}_2 \rightarrow \)
2: \( \text{Fe} + \text{O}_2 \rightarrow \)
3: \( \text{P} + \text{O}_2 \rightarrow \)
4: \( \text{N}_2 + \text{H}_2 \rightarrow \)
5: \( \text{Al} + \text{NaOH} \rightarrow \)
6: \( \text{Fe} + \text{H}_2\text{O} \rightarrow \)
7: \( \text{CH}_4 + \text{O}_2 \rightarrow \)
8: \( \text{KClO}_3 \rightarrow \)
9: \( \text{Cu} + \text{H}_2\text{SO}_4 \text{ conc.} \rightarrow \)
10: \( \text{Na}_2\text{SO}_4 + \text{H}_2\text{SO}_4 \rightarrow \)
11: \( \text{CaCO}_3 + \text{HCl} \rightarrow \)
12: \( \text{AgNO}_3 + \text{HCl} \rightarrow \)
13: \( \text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \)
14: \( \text{K}_2\text{CrO}_4 + \text{PbCl}_2 \rightarrow \)
15: \( \text{As(NO}_3)_3 + \text{H}_2\text{S} \rightarrow \)
16: \( \text{ZnCO}_3 + \text{H}_2\text{S} \rightarrow \)
17: \( \text{PbSO}_4 + (\text{NH}_4)_2\text{S} \rightarrow \)
18: \( \text{CaO} + \text{O}_2 \rightarrow \)
19: \( \text{H}_2\text{O} \text{ electrolysis} \rightarrow \)
20: \( \text{Hg} + \text{HCl} \rightarrow \)
21: \( \text{Fe}_2\text{O}_3 + \text{Al} \rightarrow \)
22: \( \text{AgNO}_3 + \text{H}_2\text{S} \rightarrow \)
23: \( \text{CuO} + \text{H}_2 \rightarrow \)
24: \( \text{Fe}_2\text{O}_3 + \text{C} \rightarrow \)
25: \( \text{K}_2\text{Fe(CN)}_6 + \text{PbCl}_2 \rightarrow \)
26: \( \text{K}_2\text{Fe(CN)}_6 + \text{PbCl}_2 \rightarrow \)
27: \( \text{K} + \text{H}_2\text{O} \rightarrow \)
28: \( \text{Al}_2(\text{SO}_4)_3 + \text{NH}_4\text{OH} \rightarrow \)
29: \( (\text{NH}_4)_2\text{SO}_4 + \text{KOH} \rightarrow \)
30: \( \text{NaNO}_3 + \text{H}_2\text{SO}_4 \rightarrow \)
31: \( \text{MnO}_2 + \text{HCl} \rightarrow \)
32: \( \text{Cu} + \text{Cl}_2 \rightarrow \)
33: \( \text{SbCl}_3 + \text{H}_2\text{S} \rightarrow \)
34: \( \text{Pb(NO}_3)_2 + \text{H}_2\text{S} \rightarrow \)
35: \( \text{CaC}_2 + \text{H}_2\text{O} \rightarrow \)
36: \( \text{Cr}_2\text{O}_3 + \text{Al} \rightarrow \)
37: \( \text{Cu} + \text{H}_2\text{O} \rightarrow \)
38: \( \text{NaI} + \text{Br}_2 \rightarrow \)
39: \( \text{Na}_2\text{CO}_3 + \text{AgNO}_3 \rightarrow \)
40: \( \text{NaCl} + \text{AgNO}_3 \rightarrow \)
41: \( \text{Sb} + \text{Cl}_2 \rightarrow \)
42: \( \text{H}_2\text{SO}_4 \text{ dil.} + \text{Fe} \rightarrow \)
43: \( \text{Al} + \text{HCl} \rightarrow \)
44: \( \text{CaCO}_3 + \text{HCl} \rightarrow \)
45: \( \text{H}_2\text{SO}_4 + \text{Ca(OH)}_2 \rightarrow \)
46: \( \text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \)
47: \( \text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 \rightarrow \)
48: \( \text{CaO} + \text{H}_2\text{O} \rightarrow \)
49: \( \text{NaCl} + \text{HNO}_3 \rightarrow \)
50: \( \text{Fe}_2\text{O}_3 + \text{CO} \rightarrow \)
51: \( \text{K} + \text{HCl} \rightarrow \)
52: \( \text{Ca} + \text{H}_2\text{O} \rightarrow \)
53: \( \text{CaSO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \)
54: \( \text{NaNO}_3 + \text{KCl} \text{ (heated)} \rightarrow \)
55: \( \text{MgBr}_2 + \text{Cl}_2 \rightarrow \)
56: \( \text{NaHCO}_3 + \text{HCl} \rightarrow \)
57: \( \text{KHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \)
58: \( \text{Na}_2\text{H}_2\text{O} + \text{H}_2\text{SO}_4 \rightarrow \)
59: \( \text{KCl} + \text{H}_2\text{SO}_4 \rightarrow \)
60: \( \text{Cu} + \text{conc.} \text{ HNO}_3 \rightarrow \)
61: \( \text{Cu} + \text{conc.} \text{ H}_2\text{SO}_4 \rightarrow \)
62: \( \text{Fe}_2\text{O}_3 + \text{Al} \rightarrow \)
63: \( \text{Cu} + \text{dil.} \text{ HNO}_3 \rightarrow \)
64: \( \text{Ca} + \text{(OH)}_2 \rightarrow \text{H}_2\text{SO}_4 \rightarrow \)
65: \( \text{Al} + \text{O}_3 \rightarrow \)
66: \( \text{CuSO}_4 + \text{H}_2\text{S} \rightarrow \)
67: \( \text{Zn(NO}_3)_2 + \text{H}_2\text{S} \rightarrow \)
68: \( \text{HgO} \text{ heated} \rightarrow \)
69: \( \text{CaCO}_3 + \text{H}_2\text{O} \rightarrow \)
70: \( \text{Ca} + \text{(HCO}_3)_2 \rightarrow \)
71: \( \text{Na}_2\text{SO}_4 + \text{HCl} \rightarrow \)
72: \( \text{NH}_4\text{HCO}_3 + \text{NaCl} \rightarrow \)
73: \( \text{NH}_3 + \text{H}_2\text{O} \rightarrow \)
74: \( \text{NH}_3 + \text{HCl} \rightarrow \)
75: \( \text{SO}_4 + \text{H}_2\text{O} \rightarrow \)
GENERAL RULES OF SOLUBILITY

All nitrates, acetates, and chlorates are soluble.

All common compounds of sodium, potassium and ammonium are soluble.

All chlorides are soluble except silver, lead and mercurous mercury. (Lead chloride dissolves considerably in hot water.)

All sulfates are soluble except lead, barium, strontium, and calcium.

The normal carbonates, phosphates, silicates and sulfides are insoluble except those of sodium, potassium and ammonium.

Hydroxides are insoluble except sodium, potassium, ammonium, calcium, barium, and strontium.
There are four major types of reactions:

**ANALYSIS**

**SYNTHESIS**

**SINGLE REPLACEMENT**

**DOUBLE REPLACEMENT**

1. Collect the necessary materials:

   You are asked to perform two experiments of each of the types listed above. Identify all products of reaction. You will have to consult your own textbook and auxiliary texts to find these tests.

2. Perform the experiment:

   Report the methods you applied and give complete balanced equations for every reaction used including those in testing for ions or elements or radicals.

---

**REACTIONS THAT GO TO COMPLETION**

1. Collect the necessary materials:

   You are to prepare and perform two reactions that go to completion by the formation of a gas; two because a precipitate is formed; and two because a slightly ionized substance results.

2. Perform the experiment:

   Report your work with ionic equations for each of these reactions and clearly indicate which ions are removed from the field of action, and how.

---

**ELECTROMOTIVE SERIES**

1. Collect the necessary materials:

   | Na metal | HCl \(\text{H}_2\text{O}_2\) dilute |
   | Ca metal | HCl dilute and concentrated |
   | Mg      | \(\text{H}_2\text{SO}_4\) dilute and concentrated |
   | Al      | \(\text{HNO}_3\) dilute and concentrated |
   | Cu      | glass plates |
   | Zn      | splints |
   | Fe      | litmus |
   | Hg      | test tubes |
   | Sn      | burner |
   | Cd      | flame test wire |
   | Pb      | |

---
Electromotive series continued:

2: Perform the experiment:
Test the reaction of Na and Ca with H\textsubscript{2}O, using a minute quantity of metal in a covered beaker. Test the products of the reaction.

Using the suggested acids and metals, determine the relative ease with which a metal will replace hydrogen from an acid; compare the acids as well as the metals. Use dilute HCl as the standard acid against which to check metal reactivity. Test all products.

3: Report the experiment in the following fashion:
   a. Arrange the metals in the descending order of reactivity.
   b. Arrange the acids in the descending order of reactivity.
   c. Balance equations for every reaction that occurred.
   d. Write out a more complete EMF Series from your text and check off in red those elements which release H\textsubscript{2} from water; in blue those which release H\textsubscript{2} from acids; in yellow those which do not release H\textsubscript{2} at all.

SALTS

General Methods of Preparation

1. Metal and acid.
2. Metallic oxide and acid.
3. Carbonate and acid.
4. Neutralization of an acid by a base.
5. Direct combination (for binary salts)

FOR SOLUBLE SALTS

General methods of preparation

1. Salt and an acid.
2. Reaction between two salts.
3. Oxide and an acid.

FOR INSOLUBLE SALTS

Review your table of solubilities and prepare two soluble and two insoluble salts.

NOTE: To prepare pure soluble salts from soluble salts, the intermediate stage must be insoluble; and conversely, to prepare an insoluble salt from an insoluble salt, the intermediate stage must be soluble.

To go from soluble to insoluble, direct transformations are possible; one can also go directly from insoluble to soluble.

Prepare NaCl by at least four different procedures.
Prepare BaSO\textsubscript{4} by at least four different procedures.
Give detailed explanation, with equations, of every step you use. Report this experiment by a flow sheet of equations.
Name of school: 

Location of school: 

Name of companion-school teaching the Unit inductively (or deductively) 

Name of teacher: 

Number of years experience as chemistry teacher: 

Number of pupils present for pre-test: For post-test: 

Grade in which this unit is being taught: 

Number of boys in the class: Girls: 

Do you teach macro or semimicro laboratory work? (Circle one) 

What units have you taught so far in chemistry this year? 

What textbook does your class use? 

by: pub.by: 

Number in your class who have had a course in: 

- General Science: Biology: Physics:
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You will see a test sheet of pictures representing chemical experiences. The first eight questions may be answered on the following blanks. In the reactions, A produces B. The reasons for each reaction are given in the picture. On your answer check:

a) If the reason given is true and applies to the reaction.
b) If the reason given is true and does not apply to the reaction.
c) If the reason given is false.

1: 
2: 
3: 
4: 
5: 
6: 
7: 
8: 

In the following reactions, the volume of the products is related to the volume of the reactants so that:

a) An increase in A produces an increase in B.
b) An increase in A produces a decrease in B.
c) An increase in A produces no change in B.

9: 
10: 
11: 
12: 
13: 
14: 
15: 
16: 
The following reactions are:
a) Correctly represented and go to completion.
b) Correctly represented and do not go to completion.
c) Incorrectly represented.

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In the following reactions:
a) An increase in A is accompanied by an increase in B.
b) An increase in A is accompanied by a decrease in B.
c) An increase in A does not affect B.

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The following reactions:
a) Observe the Law of Conservation of Matter and are correctly balanced.
b) Do not observe the Law of Conservation of Matter and are correctly balanced.
c) Do not obey the Law of Conservation of Matter; are incorrectly balanced.

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In regard to the following reactions, they
a) will react and yield the products indicated.
b) will react and will not yield the products indicated.
c) will not react.

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1. Hydrogen is the lightest gas known.

2. Calcium is an active metal.

3. H₂S forms an insoluble salt with As⁺⁺⁺.

4. MnO₂ contains oxygen.

5. The reaction is complete.

6. The reaction is complete.

7. CuO is a precipitate.

8. Oxygen is a gas.

9. The reaction is complete.

10. Nickel chloride is insoluble.

11. SbCl₅ is dissolved in water.
Increase volume of carbon monoxide

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase temperature

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase amount of catalyst

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase temperature

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase pressure

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase pressure

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase volume of carbon monoxide

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase amount of catalyst

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Increase temperature

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$

Increase pressure

$2 \text{H}_2 + \text{CO} \rightarrow \text{CH}_3\text{OH} + \Delta$
21. 

Fe

amount of catalyst

Zn + H₂SO₄

vol. H₂

22. 

ACTIVITY OF A

EASE OF REDUCTION OF THE

METAL

METALLIC OXIDE

23. 

3 H₂ + N₂ → 2 NH₃

6 liters + 2 liters → 6 liters

24. 

carbon monoxide + oxygen → carbon dioxide

48 liters + 24 liters → 48 liters

25. 

6 liters

26. 

AgNO₃ + NaCl → 342

AgCl ↓ + NaNO₃

27. 

H₂SO₄ + Cd → CdSO₄ + H₂

98 gr. + 115 gr. = 209 gr. + 2 gr.
31. $\text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$

32. $\text{Pb(NO}_3\text{)_2} + \text{H}_2\text{SO}_4 \rightarrow \text{PbSO}_4 + \text{H}_2\text{NO}_3$

33. $\text{conc. H}_2\text{SO}_4 + \text{Pb} \rightarrow \text{PbSO}_4 + \text{H}_2$

34. $\text{dil. H}_2\text{SO}_4 + \text{Cu shot} \rightarrow \text{H}_2\text{SO}_4 + \text{Cu shot}_e$

35. $2 \text{Na} + \text{H}_2\text{O} \rightarrow 2 \text{NaO} + \text{H}_2$

36. $\text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$

37. $\text{aluminum} + \text{H}_2\text{O} \rightarrow \text{H}_2$

38. $\text{aluminum nitrate} + \text{dil. H}_2\text{SO}_4 \rightarrow \text{H}_2$

39. $\text{conce. HNO}_3 + \text{Cu} \rightarrow 3 \text{Cu(NO}_3\text{)_2} + 2 \text{NO}_3 + 4 \text{H}_2\text{O}$

40. $\text{Al} + 5 \text{HNO}_3 \rightarrow \text{Al(NO}_3\text{)_3} + \text{H}_2$
The traditional method of teaching high school chemistry in America is what we are designating as the DEDUCTIVE-DEDUCTIVE method. In our study we are contrasting this with the inductive method, hence, we want all the classes TAUGHT deductively to be taught in a THOROUGHLY DEDUCTIVE FASHION. In order to insure this we are presenting you with an outline of classroom lectures for the unit on chemical equation-balancing. In addition, in order to insure more complete uniformity among the classes, we are including sheets of laboratory directions for each student. Will you please use these for the experiments of this unit rather than those of the regular manual you are following this year?

The pre- and post-tests you have here are the same as those given to the inductive classes and are designed to measure understanding of the concepts involved in equation-balancing.

As an over-all survey of the material to be covered, here is an outline of the unit:

I: WORD REACTIONS:
   a) state experimental facts accurately
   b) write the reactants on the left side of the arrow; write products on the right.

II: SYMBOLIZED EQUATION from the word reaction:
   a) write the correct formulas for all reactants and products.

III: BALANCED EQUATIONS:
   a) Law of Conservation of Matter
   b) make the total number of atoms on the left of the arrow equal the total number on the right, using the smallest possible common denominator as balancing number.

IV: BALANCING OF EQUATIONS:
   a) the reactants and products of the equation
   b) the composition of the molecules that react
   c) the number of each kind of molecule taking part
   d) the weights of reactants and products
   e) the volumes of the gaseous substances

V: CLASSIFICATION OF REACTIONS:
   a) direct combination or synthesis
   b) simple decomposition or analysis
   c) single or single replacement or substitution
   d) double replacement, double decomposition, or metathesis

VI: REACTIONS INVOLVING SOLID-STATE EQUILIBRIUM:
   a) formation of a gas
   b) formation of a precipitate; rules for solubility
   c) formation of a slightly ionized substance

VII: EQUILIBRIUM:
   a) Law of Mass Action
   b) Le Chatelier's Principle
We would like your evaluation of the success of this method and for this purpose we are providing several headings and are asking you to write comments as indicated in the samples.

1. The skill and knowledge shown in laboratory work, aside from the accuracy of testing and identification, can evaluate general ability and intelligence. Sample:

   Sally J. has shown a systematic approach to the task and is careful and precise. Please fill in your comments in the space given.

2. The interest shown in laboratory work, the attentive nature, observations, questions, and inventiveness, all are indications of the student's acquisition of a scientific attitude. Sample:

   John W. is in the lab to try out an experiment with Kilo. He is attentive and curious.

3. The problem-solving part of the unit in which the student must devise, carry out, and explain laboratory procedures is a good indication of the student's ability to reason. Sample:

   Stephen T. has a well-organized notebook with detailed observations, conclusions, and references.

4. The number and kinds of chemicals and equipment which the student requests are a good indication of his growth in and assimilation of chemical knowledge. Sample:

   Doug P. requested additional metals to further test the EPK hypotheses.

5. Use of reference to assist in the solution of difficulties that arise in experimentation indicates an ability to coordinate the work in hand. Sample:

   Lance P. turned to the "Handbook of Chemistry and Physics" to determine the relative solubilities of 2 and 3 ions.
The type of reference book used and the student interpretation of it are clear indications of his level of learning.

Sample:

ﻇیح یاهو یاهو فیولولو لکی فیولولو فیولولو لکی

The organization of the final report sheets will show the degree to which the material has been assimilated.

Sample:

James W. filed in untidy, unorganized report sheet

The procedures used in the laboratory will differentiate between leaders and followers, thinkers and technicians. Students who merely follow the verbatim directions found in a text will not be classed as leaders.

Sample:

Betsy is not to obey the directions, do an experiment, then she

We are asking you for intelligence test scores on the pupils being used in this study. Will you please label each one, and all pre- and post-tests with the first name and last initial of the pupil? If IQ's are not available in your school and you would be willing to give an Otis Quick Scoring Test, will you notify me immediately by means of the enclosed card which you return when the packet arrives? I will gladly send you test booklets and answer sheets, and will score them if you wish. If, however, your school authorities do not wish IQ's revealed, send us your data anyway as they will still be of great value.

We have included a school data sheet which we ask you to fill out. This along with several other sheets, is stamped "please return." Will you send these papers to the teacher of the inductive unit from whom you received this material so that he, in turn, may return the complete packet to me.

We hope that you will enjoy working this unit with us and we wish you the best of luck with your classes this year.
CLASS I: WORD REACTIONS

A series of introductory reactions should be discussed, as: coal + air = ashes + water-vapor + carbon dioxide; iron + moist air = rust. Students may volunteer familiar reactions until ten or a dozen have been accumulated. Reactants and products are then labelled, identified, and described. Definitions of both are given.

The Law of Conservation of Matter and the Law of Conservation of Energy should be stated: Neither matter nor energy can be created nor destroyed in any chemical reaction. Some time should be given for the memorization of this law.

Point out that the products and reactants of any reaction are equal in weight, even though they appear at first glance not to be.

CLASS II: SYMBOLIZED EQUATIONS; SYNTHESIS REACTIONS

The alkali metal, alkaline earth and the halogen families should be placed on the board by symbol and the valence of all should be marked. Write also a list of common radicals. (These lists are included in this unit for your convenience.) Spend some time in reviewing student ability to write compounds by properly combining the symbols for metals with those for non-metals and radicals.

Synthesis is defined as an action in which a more complex substance is made from two or more simpler ones. Examples are written, such as: C + O₂ → CO₂; S + O₂ → SO₂; H₂ + Cl₂ → 2 HCl; 2 Zn + O₂ → 2 ZnO; Pb + S → PbS; 2 H₂ + O₂ → 2 H₂O. Point out that the total number of atoms to the right of the arrow must equal the total number to the left of the arrow in order to comply with the Law of Conservation of Matter. In each equation point out:

A: reactants and products; are they correctly written?
B: composition of the molecules; are the formulas correct?
C: the number of each kind of molecule taking part in the reaction
D: weights of reactants and products; are they equal?
E: volumes of gaseous substances; are they correctly represented?

CLASS III: ANALYSIS REACTIONS

List at least five of these and proceed to explain them as in the case of the synthesis reactions, using the points A-E as listed above. Include the following reactions:

2 KClO₃ → 2 KCl + 3 O₂
2 H₂O → 2 H₂ + O₂
2 H₂O → 2 H₂ + O₂
Repeat steps A-E as before, illustrating by typical replacement reactions such as:

\[ \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]
\[ \text{Fe} + \text{H}_2\text{SO}_4 \rightarrow \text{FeSO}_4 + \text{H}_2 \]
\[ 2 \text{Na} + 2 \text{H}_2\text{O} \rightarrow 2 \text{NaOH} + \text{H}_2 \]
\[ \text{Ca} + \text{MnCl}_2 \rightarrow \text{CaCl}_2 + \text{Mn} \]
\[ \text{Cu} + \text{conc. HNO}_3 \rightarrow \text{no hydrogen} \]
\[ \text{Cu} + \text{HCl} \rightarrow \text{no hydrogen} \]

The electromotive series must then be explained as an arrangement of elements in the order of reactivity. The more active metals replace the less active ones from the compounds in which they are found. A device for remembering the EMF series may be given:

Little Karen naturally craved an MG, and Mandy's zinc car, fearing Connie's nine specky plumbers. or: Little Kathy Nash caught midgies all morning. Zane Cadwell fed cousin Dick's son's plumber.

CLASS VI: DOUBLE DISPLACEMENT REACTIONS; CLASSIFICATION OF REACTIONS

Repeat steps A-E with this type of reaction. Illustrate with common examples:

\[ \text{AgNO}_3 + \text{NaCl} \rightarrow \]
\[ \text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \]
\[ \text{NaOH} + \text{HNO}_3 \rightarrow \]

Assign the memorization of the solubility charts. (These are included in this unit.)

Students must now learn to recognize analysis, synthesis, double decomposition, and single replacement reactions when they meet them. Have some students work at the board balancing and classifying reactions that you dictate.

CLASS VI: REACTIONS THAT GO TO COMPLETION

Illustrate each type of reaction:

\[ \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]
\[ \text{AgNO}_3 + \text{HCl} \rightarrow \text{HNO}_3 + \text{AgCl} \]
\[ \text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]

The above reactions may then be written ionically:

\[ \text{Zn}^2+ + 2 \text{H}^+ + 2 \text{Cl}^- \rightarrow \text{Zn}^{++} + 2 \text{Cl}^- + \text{H}_2 \]
\[ \text{Ag}^+ + \text{NO}_3^- + \text{H}^+ + \text{Cl}^- \rightarrow \text{H}^+ + \text{NO}_3^- + \text{AgCl} \]
\[ \text{Na}^+ + \text{OH}^- + \text{H}^+ + \text{Cl}^- \rightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O} \]

The removal of ions from the field of action permits a reaction to go to completion.

Review the solubility table by oral quizzing. Dictate various equations from the list in this section and ask the type of reaction and the reason it goes to completion.
Reactions that do not go to completion start reacting in reverse and we have two reactions going on simultaneously with products and reactants interchanged.

\[ \text{N}_2 + 3 \text{H}_2 \rightleftharpoons 2 \text{NH}_3 \]
\[ 2 \text{NH}_3 \rightleftharpoons \text{N}_2 + 3 \text{H}_2 \]

No ions are removed from the field of action and hence both reactions occur. Variations in concentration, in temperature, and in pressure can affect the equilibrium point so that one of the two reactions is favored. If the products of one reaction are removed, then that reaction will proceed and the reverse reaction will be minimized.

Analyze the reaction: \( \text{NaNO}_3 + \text{KCl} \rightleftharpoons \text{KNO}_3 + \text{NaCl} \)
According to the table of solubilities, all these salts are very soluble at room temperature, hence, the reaction cannot go to completion because of the insolubility of a product. Again, none of the salts is volatile, so, volatility of product is no cause of completion. Moreover, water is not formed; nor is any other slightly ionized substance. Written ionically the reaction is:

\[ \text{Na}^+ + \text{NO}_3^- \rightleftharpoons \text{K}^+ + \text{Cl}^- \rightleftharpoons \text{Na}^+ + \text{Cl}^- + \text{K}^+ + \text{NO}_3^- \]

Stress the fact that equilibrium is a dynamic, not a static state. State the Law of Mass Action: The rate of a chemical reaction is proportional to the concentration of the reacting substances, other factors being unchanged.

CLASS VIII: LE CHATELIER'S PRINCIPLE

If a stress or strain is applied to a system in equilibrium, that reaction occurs which tends to offset the stress or strain.

Removal of a product from the field of action is the only positive and permanent way to break an equilibrium.

\[ \text{N}_2 + 3 \text{H}_2 \rightleftharpoons 2 \text{NH}_3 \]
\[ 1 \text{vol.} + 3 \text{vol.} \rightarrow 2 \text{vol.} \]
\[ 4 \text{vol.} \rightarrow 2 \text{vol.} \]

Larger volumes are associated with high temperature, therefore, increased temperature will favor the decomposition of ammonia, the reverse reaction as written above. Smaller volumes are associated with high pressure, therefore, an increase of pressure on the system will favor the forward reaction.

Catalysts may change the equilibrium point temporarily. Eventually all reactions in equilibrium tend to return to a state of equilibrium once the stress or strain applied is offset.
10-40 minute lab. sessions or 5 double periods

DEDUCTIVE LABORATORY NOTES

-1-

TYPES OF CHEMICAL REACTION

I: COMBINATION or SYNTHESIS:

1: On a piece of paper place a spoonful of powdered sulfur and a spoonful of iron filings. Take a pinch of the sulfur and dissolve it in 2 ml. of carbon disulfide; bring a magnet near the iron filings and note the way in which they are attracted to it.

2: In an old test tube mix one gram of sulfur with one gram of iron filings and heat until the contents of the tube become bright cherry-red in color. Plunge the red hot tube into cold water to crack the glass; let the chunk of iron sulfide cool before you touch it, then try to dissolve it in carbon disulfide; touch it with a magnet to determine if it is magnetic.

3: Using a pair of forceps, heat about three inches of magnesium ribbon in the flame. Do not look directly at the burning magnesium. The white powder formed is magnesium oxide.

II: DECOMPOSITION or ANALYSIS:

1: Place about a gram of mercuric oxide in a small test tube and heat the tube in your burner. Test the escaping gas with a glowing splint. Notice the deposition of mercury globules on the walls of the tube.
2: Place a few marble chips in a test tube fitted with a delivery tube which extends into another test tube full of calcium hydroxide. Heat the marble chips and note that they decompose producing the gas carbon dioxide which turns the calcium hydroxide milky. This is a standard test for the identification of carbon dioxide.

III: SINGLE REPLACEMENT: Hydrogen is displaced from an acid by a metal

1: Take three test tubes and into the first put a portion of granulated zinc; into the second, iron filings, and into the third, magnesium powder. Add 2 ml. HCl to each. Test the escaping gas with a burning splint. A characteristic "pop" is the common test for hydrogen gas.

IV: DOUBLE DECOMPOSITION or METATHESIS:

1: To 3 ml. dil. HCl in a test tube, add 1 ml. silver nitrate. This reaction is the test for the presence of the chloride ion.

2: To 3 ml. lead nitrate in a test tube, add 1 ml. potassium chromate. The yellow precipitate is a confirmatory test for lead.

3: Place a few crystals of calcium carbonate (marble chips) in a test tube and add 2 ml. dil. HCl. The fizzing action that results from the escape of carbon dioxide is the test geologists use to detect a carbonate.

4: To 3 ml. sodium sulfate in a test tube, add 1 ml. barium chloride. The formation of a white precipitate which does not dissolve in dil. HCl is a positive test for a sulfate.
REPORT: Report these reactions by defining the type of reaction and by placing balanced equations for the reactions performed under the proper definitions.

ELEcTRoCHEMICAL SERIES

Problem: To determine the relative reactivities of common metals.

Procedure: Set six test tubes in a rack and half fill each with water. Add 1 ml. conc. HCL to each. To the first, add a piece of mossy zinc; to the second, a piece of copper shot; to the third, an iron nail; to the fourth, a piece of lead, and to the fifth a strip of magnesium ribbon. In the last, place a piece of tin. Note the reaction in each case. Test the escaping gas in each tube with a lighted splint.

2: Set up 5 test tubes in your rack. To the first, add a solution of silver nitrate; to the second, magnesium sulfate; to the third, zinc sulfate; to the fourth, copper sulfate, and to the fifth, sodium chloride. Clean a strip of magnesium ribbon with sandpaper and immediately place a 2" piece in each and carefully note the effect. Repeat the set-up using strips of Zn (cleaned in the same way) in place of the magnesium.

REPORT: Using the following table, write the reactions which occurred.

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<tr>
<th>METAL</th>
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<tr>
<td>Aluminum</td>
<td>$2 \text{Al} + 6 \text{HCL} \rightarrow 2 \text{AlCl}_3 + 3\text{H}_2$</td>
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<td>Magnesium</td>
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-3-

WHY REACTIONS GO TO COMPLETION

Problem: To determine whether or not a chemical reaction has

Procedure: To 3 ml. of sodium chloride solution in a test tube, add 1 ml. silver nitrate. A precipitate is formed.

To 3 ml. of sodium sulfate in a test tube, add 1 ml. of barium chloride. Again a precipitate is formed.

In both of these cases, ions are removed from solution by the formation of an insoluble compound, and hence, the reaction goes to completion.

2: FORMATION OF GASEOUS PRODUCTS: To a little NaCl in a test tube, add 1 ml. dil. sulfuric acid and test the escaping gas with blue litmus paper that has been moistened in water.

To some solid ammonium chloride in a test tube, add some sodium hydroxide solution and heat gently. Test the escaping gas with moist red litmus which will turn blue in acid.

3: FORMATION OF A SLIGHTLY IONIZED PRODUCT:

Pour about 10 ml. sodium hydroxide solution into a beaker. Neutralize it with dilute HCl until litmus paper shows neutrality. Taste the resulting solution. Evaporate the solution to dryness and note the formation of salt crystals as the water escapes.
REVERSIBLE REACTIONS

Purpose: To study the conditions needed for equilibrium.

Procedure: Half fill a dry test tube with alcohol and put the same amount of water into another test tube. Add a piece of cupric bromide about the size of a pea to both tubes. You will see the characteristic color of cupric ions in the test tube containing the water solvent. Cupric bromide does not ionize in alcohol solution.

Pour the water solution of ionized cupric bromide into a small beaker and add an equal volume of sodium chloride solution. You will find no evidence of reaction because here we have an equilibrium. Write the reversible reaction as you have already learned to do in class, and show what four ions are present as reactants and products.

PREPARATION OF SALTS

Purpose: To determine how salts can be prepared.

Procedure: The first method to be employed is

1: DIRECT COMBINATION of a metal and non-metal to produce a binary salt. Place 3 grams of potassium chlorate in a clean dry gas bottle or large Erlenmeyer flask, and add to it 2 ml. conc. HCl. Cover the bottle with a glass place and wait until the green fumes of chlorine have filled the bottle. Slowly add powdered antimony by sprinkling the metal into the chlorine. White fumes of antimony chloride will form and these can be hydrolyzed by water until they settle out.

2: SINGLE REPLACEMENT: Place a strip of magnesium ribbon in a test tube and add 1 ml. conc. HCl. Test the gas that escapes by the aid of a burning splint. Allow the reac-
tion to proceed until all the metal has been used up. (You may have to add a bit more acid to dissolve the magnesium totally.) Evaporate the resulting solution until the salt, magnesium chloride remains.

To 3 ml. conc. nitric acid in a test tube, add a few pieces of copper shot. Note the color of the gas evolved. Copper never frees hydrogen from an acid, but does react with nitric and concentrated sulfuric acids to release oxidation products. This brown gas is nitrogen dioxide. Heat the solution resulting from the action of the acid and copper until no liquid remains. Copper salts are always characterized by colors such as you have in this salt.

3: DOUBLE REPLACEMENT: In a test tube place one gram of cupric oxide and add to it 2 ml. conc. nitric acid. Heat the tube gently. Dilute the solution by filling the test tube with water and note the resulting color. Evaporate the solution to dryness by transferring the liquid to an evaporating dish and heating it gently.

Dissolve a gram of sodium chloride in 5 ml. of water and add 1 ml. of silver nitrate solution. A white precipitate forms. Let this precipitate stand in the direct sunlight or under an electric light bulb for a few minutes and note the change in the color of it. Sunlight always affects these silver salts, and hence they are useful in photography.

Place 3 ml. sodium carbonate solution in a test tube and add dil. HCl gradually until there is no more effervescence. When the litmus paper shows the solution to be neutral, add one drop of HCl in excess. Avoid adding a large excess of acid. Evaporate the solution to dryness and taste the residue.

To 2 ml. dil. sodium hydroxide in a test tube, add dil. HCl until the solution is neutral to litmus. Evaporate the water off and taste the resulting solid.
APPENDIX G

TABLES OF SCHOOL DATA - PHASE II
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\sum y^2 = 7993 \\
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\[ \Sigma x = 270 \]
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Type A School XI

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\[
\begin{align*}
\Sigma x &= 291 \\
\Sigma y &= 395 \\
\Sigma x^2 &= 3576 \\
\Sigma y^2 &= 6917 \\
\Sigma xy &= 4535 \\
\Sigma x &= 205 \\
\Sigma y &= 482 \\
\Sigma x^2 &= 3671 \\
\Sigma y^2 &= 9854 \\
\Sigma xy &= 5717
\end{align*}
\]
### Type A School XII

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\sum X = 231 \quad \sum X^2 = 3893 \\
\sum Y = 515 \quad \sum Y^2 = 5047 \\
\sum XY = 5775
\]

\[
\sum X = 237 \quad \sum X^2 = 5095 \\
\sum Y = 402 \quad \sum Y^2 = 8234 \\
\sum XY = 4737
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| Σ x² | 435 | Σ x² | 4923 |
| Σ xy | 5150 | Σ xy | 7772 |
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| ΣXY  | 3187| ΣXY²| 9426 |

| ΣX   | 267 |
| ΣY   | 666 |
| ΣXY  | 3187|
Type A School XV

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\Sigma x'^2 &= 3906 \\
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\[
\begin{align*}
\sum x &= 272 \\
\sum y &= 415 \\
\sum x^2 &= 5682 \\
\sum y^2 &= 8335 \\
\sum xy &= 5444 \\
\sum x^2 &= 350 \\
\sum y^2 &= 456 \\
\sum x^2 &= 4686 \\
\sum y^2 &= 7313 \\
\sum xy &= 5585
\end{align*}
\]
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\[ \sum y^2 = 9183 \]
\[ \sum xy = 5596 \]

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\[ \sum x^2 = 4264 \]
\[ \sum y^2 = 5996 \]
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\[
\Sigma x = 297 \\
\Sigma y = 572 \\
\Sigma x^2 = 4015 \\
\Sigma y^2 = 5554 \\
\Sigma xy = 4721
\]

\[
\Sigma i = 353 \\
\Sigma y = 591 \\
\Sigma y^2 = 4699 \\
\Sigma y^2 = 11921 \\
\Sigma xy = 6765
\]
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\[
\begin{align*}
\sum x &= 294 \\
\sum y &= 369 \\
\sum x^2 &= 3972 \\
\sum y^2 &= 5557 \\
\sum xy &= 4534 \\
\end{align*}
\]

\[
\begin{align*}
\sum x &= 257 \\
\sum y &= 498 \\
\sum x^2 &= 2841 \\
\sum y^2 &= 9000 \\
\sum xy &= 5206 \\
\end{align*}
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\[ \sum x^2 = 3516 \]
\[ \sum y^2 = 12454 \]

\[ \sum xy = 6349 \]

\[ \sum x = 338 \]
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\[ \sum x^2 = 4184 \]
\[ \sum y^2 = 6326 \]

\[ \sum xy = 5114 \]
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\Sigma x &= 256 \\
\Sigma y &= 358 \\
\sum x^2 &= 4127 \\
\sum y^2 &= 4567 \\
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Type B School VII

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\[ \sum X^2 = 2015 \quad \sum Y^2 = 10426 \]
\[ \sum XY = 5928 \quad \sum X^2Y = 2355 \quad \sum XY^2 = 4025 \quad \sum XY = 2935 \]
### Type D School Totals

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Type B School VIII Late data, not included in Chapter VII

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A \(t\)-value of 2.01 is significant at the .05 level of confidence. Hence, despite the handicap of lower intelligence, the inductive group made a significantly greater gain than did the deductive group on this unit.
BIBLIOGRAPHY


TEST BIBLIOGRAPHY


2. California Test of Mental Maturity, Advanced Form, California Test Bureau, Los Angeles, California, 1946.
